GENERAL CHEMISTRY II LAB MANUAL



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General Chemistry II CHM 141 Lab Manual

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Glossary

Detailed Licensing

Detailed Licensing

Licensing

*A detailed breakdown of this resource's licensing can be found in **Back Matter/Detailed Licensing**.*

CHAPTER OVERVIEW

0: Lab Safety Orientation

0.3: Lab Safety Quiz
0.4: Additional Resources
0.1: Lab Safety Videos
0.2: Lab Safety Contract



Figure **0.1**: The four principles of safety (Copyright; ACS)

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0.3: Lab Safety Quiz

Multiple-Choice Questions

1. **What should you do if a chemical splashes into your eyes?**
   A. Rub your eyes vigorously
   B. Rinse your eyes with water for at least 15 minutes
   C. Close your eyes and wait for the pain to subside
   D. Apply a cold compress
2. **Which of the following is NOT appropriate lab attire?**
   A. Closed-toe shoes
   B. Long pants
   C. Loose, flowing clothing
   D. Safety goggles
3. **What is the first thing you should do if you spill a chemical on your skin?**
   A. Wipe it off with a cloth
   B. Rinse the affected area with water immediately
   C. Apply a bandage
   D. Ignore it if it doesn't hurt
4. **Where should you dispose of broken glassware?**
   A. Regular trash bin
   B. Recycling bin
   C. Designated glass disposal container
   D. Sink
5. **Which of the following statements about lab safety is TRUE?**
   A. Eating and drinking are allowed in the lab
   B. You should always work alone in the lab
   C. Always read labels and safety data sheets before using chemicals
   D. It's okay to leave a Bunsen burner unattended

True/False Questions

6. **You should always wear safety goggles when handling chemicals.**
   - True
   - False
6. **It's safe to taste chemicals in the lab to identify them.**
   - True
   - False
6. **You should know the location of safety equipment such as fire extinguishers, eyewash stations, and safety showers.**
   - True
   - False
6. **Long hair should be tied back when working in the lab.**
   - True
   - False
6. **It's acceptable to use your phone in the lab as long as you're careful.**
   - True
   - False

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0.4: Additional Resources



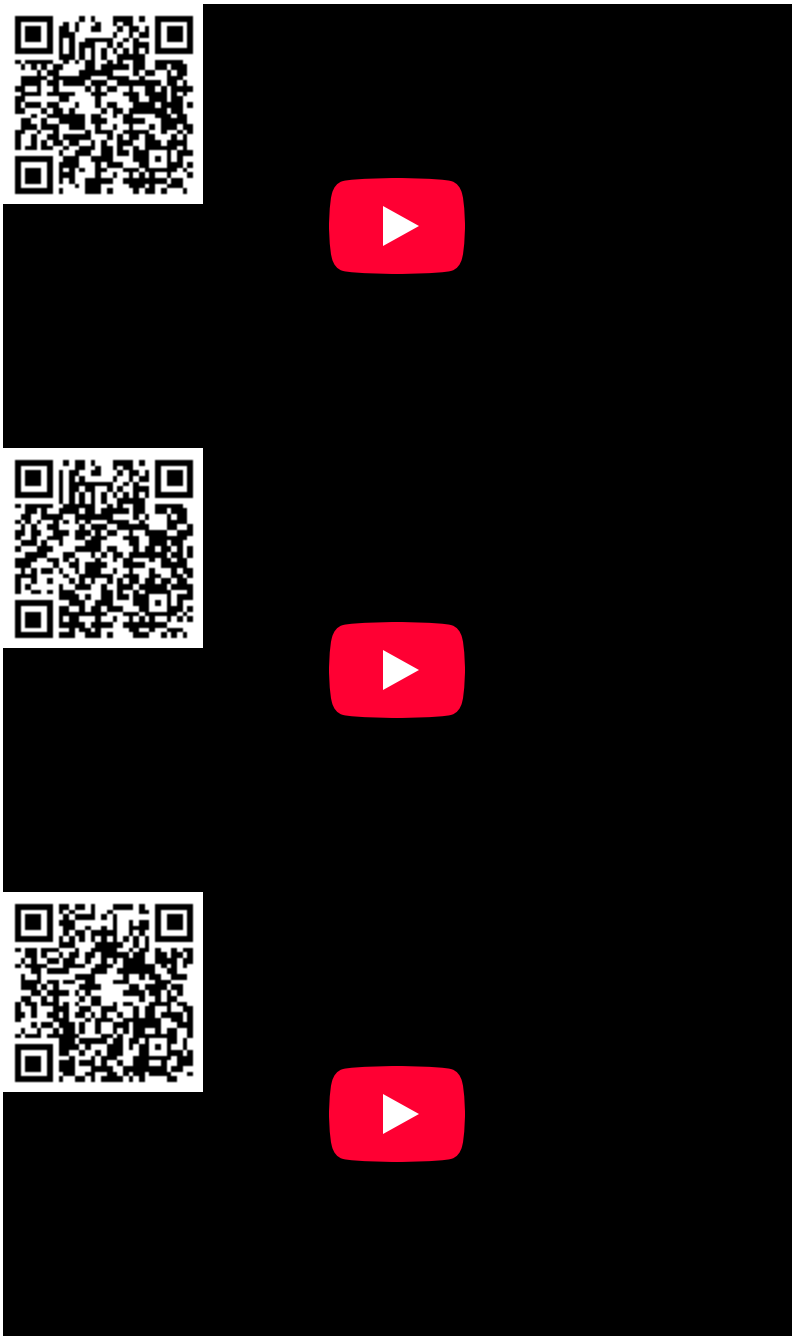
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0.1: Lab Safety Videos

Safety Orientation Videos

Students will find below the eight Lab safety orientation videos. The first six are part of the College Lab Safety Series by the American Chemical Society. Students should watch the first seven videos before signing the safety contract and attempting the safety quiz. The eighth video is an older version of the College Lab Safety video by the American Chemical Society, which should be watched, time permitting, as it provides detailed dos and don'ts in a college lab setting.

1. **ACS (American Chemical Society) College Lab Safety Series**





2. Lab Safety and Techniques- Crash course safety



3. Safety Video by American Chemical Society (1991)



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0.2: Lab Safety Contract

Triton College Chemistry Division

Student Safety Contract

At Triton College, safety is our number one priority. A list of rules has been developed and provided to you in this student safety contract to ensure a safe science laboratory experience. These rules must always be followed. Please be advised that you will be asked to leave the laboratory if you violate these rules. **No exceptions!!**

Download, sign, and date both contract pages before you participate in the laboratory sessions. Your signature indicates that you have read this contract and agree to abide by the rules. You will keep a copy for your records and upload or submit one to your instructor via your Blackboard or Learning Management System (LMS) as directed.

I, _____, dated _____ agree that I will comply with the following safety rules. I will:

1. Always wear safety goggles in the laboratory.
2. Always wear appropriate clothing in the laboratory. This includes long pants, shirts with sleeves, and closed shoes (i.e., no sandals). Also, I understand that long hair must be tied back, and loose or baggy clothing or jewelry must be secured.

   **Note: Students violating rules one and/or two will be asked to leave the lab immediately. No second chances!! NO EXCEPTIONS!!*
3. Read laboratory instructions beforehand and ask the instructor about any part of the lab I am unsure about.
4. Use proper equipment for laboratory procedures.
5. Never perform unauthorized experiments or work alone in the laboratory.
6. Never remove chemicals or other materials/equipment from the laboratory area.
7. Read the labels on reagent bottles carefully, remove only small amounts of reagent with the proper tools, and never return unused chemicals to the bottle.
8. Dispose of broken glass and waste chemicals in the appropriate waste container.
9. Immediately inform the instructor of a chemical spill or accident (even minor ones) in the laboratory.
10. Never taste or smell any chemicals.
11. Know the location of eyewash fountains, fire extinguishers, fire blankets, safety showers, first aid kits, and exits.
12. Do not bring food or drink into the lab; refrain from eating or drinking in the laboratory. This includes chewing gum and eating candy.
13. Keep aisles clear. Books, purses, backpacks, and other personal belongings will be kept outside or under the lab benches.
14. I will responsibly conduct myself at all times in the laboratory. Horseplay, practical jokes, and pranks are dangerous and prohibited.
15. Clean my work area before leaving the lab.
16. Wash my hands before I leave the lab.

Agreement

I, _____ have read and agree to follow all the safety rules outlined in this contract. I realize I must obey these rules to ensure my safety and that of my fellow students. I will fully cooperate with my instructor and fellow students to maintain a safe lab environment. I will also closely follow all oral and written instructions provided by the instructor, the science department, and the college. I am aware that any violation of this safety contract, resulting in unsafe conditions in the laboratory or misbehavior on my part, may lead to my removal from the laboratory, a failing grade, and/or dismissal from the course.

Student Signature: _____

Date: _____

Course section and number: _____

Questions

1. Are you color blind? Yes / No
2. Do you have any allergies? Yes / No
   If yes, please list specific allergies: _____
3. Do you wear contact lenses? Yes / No
4. If you are pregnant, breastfeeding, or trying to conceive, avoid working with heavy metal compounds and organic solvents. Notify your instructor and request a list of chemicals used in this course. Use this list to consult with your physician before proceeding with the course. Under your physician's supervision, you may need to purchase a respirator, which a doctor must fit and adjust to your needs.

5. Please inform your instructor of any other medical conditions or situations you feel we should be aware of to ensure your safety and the safety of others while working in the laboratory.

Please be aware that anything written in this contract will remain confidential.

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CHAPTER OVERVIEW

1: Beer's Law

PURPOSE

- To determine the concentration of an unknown copper(II) sulfate solution.
- To understand and apply Beer's Law, which relates the absorbance of a solution to its concentration.
- To prepare and measure the absorbance of standard copper(II) sulfate solutions to construct a standard calibration curve.
- To use the standard curve—or its linear regression equation—to calculate the molar concentration of the unknown solution.

INTRODUCTION

The goal of this experiment is to determine the concentration of an unknown copper(II) sulfate solution by measuring how strongly it absorbs light. In colored solutions, absorbance increases as concentration increases; a more concentrated solution absorbs more light and transmits less than a more dilute one. This predictable behavior provides the basis for quantitative analysis using visible spectroscopy.

To obtain accurate measurements, it is first necessary to identify the wavelength at which the solution absorbs most strongly. This wavelength, known as λ_{max}, provides the greatest sensitivity because small changes in concentration produce measurable changes in absorbance.

Beer's Law describes the relationship between absorbance and concentration. Once a set of standard solutions with known concentrations is prepared, their absorbance values are measured using a spectrophotometer. Each absorbance reflects the amount of light that passes through the solution and reaches a photocell inside the instrument.

Plotting absorbance versus concentration for these standard solutions produces a linear calibration curve, commonly called a Beer's Law plot or standard curve. The slope of this line captures the proportionality between absorbance and concentration and can be used to determine the concentration of an unknown solution. A strong linear fit—indicated by a correlation coefficient (r) close to 1.00 and a y-intercept (b) near zero—demonstrates reliable adherence to Beer's Law.

In this experiment, you will:

- To prepare a series of standard copper(II) sulfate solutions.
- To measure their absorbance values at λ_{max}.
- To construct a standard calibration curve using your spectrophotometer data.
- To measure the absorbance of an unknown copper(II) sulfate solution and determine its concentration using the standard curve or regression equation.

DATA PREP: THE "BEST FIT" LINE

In this lab, you will calculate the concentration of an unknown. To do this accurately, your calibration curve must be linear.

- **The Diagnostic:** We use the **Coefficient of Determination (R^2)** to grade the straightness of the line.
- **The Goal:** An R^2 of 1.000 is perfect. In this lab, if your $R^2 < 0.98$, it implies your pipetting technique was inconsistent, and your calculated unknown concentration will likely be wrong.

Key Equations

Beer's Law:

$$A = \varepsilon bc \quad \text{or} \quad A = kc \text{ (for constant } \varepsilon \text{ and path length)} \tag{1.1}$$

Dilution Equation:

$$C_1V_1 = C_2V_2 \tag{1.2}$$

1.1: Beer's Law - Experiment
1.2: Beer's Law - Pre-lab
1.3: Beer's Law - Data and Report

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1.1: Beer's Law - Experiment

SAFETY PRECAUTIONS

- Safety goggles must be worn at all times.
- All chemicals should be handled with care to avoid spills.

EQUIPMENT AND CHEMICALS NEEDED

Table **1.1.1**: Materials Needed

Equipment	Chemicals
UV-Vis spectrophotometer, cuvettes, tissues	0.10 M $CuSO_4$
Five 16x150 mm test tubes, test tube rack	$CuSO_4$, unknown solution
Two 10-mL graduated pipettes, bulb or pump	
Two 100-mL beakers	
10-mL volumetric flask	

EXPERIMENTAL PROCEDURE

Part A: Prepare Solutions

1. Get small amounts of 0.10 M $CuSO_4$ solution and deionized water in separate beakers.
2. Label five test tubes 1-5.
3. Create five standard solutions using the graduated pipettes. Measure the required amounts of Copper Sulfate and Water directly into the test tubes.

Table **1.1.2**: Standard concentrations

Test Tube	0.10 M $CuSO_4$ [mL]	Deionized water [mL]	Final $CuSO_4$ concentration [M]
1	2.0	8.0	0.020
2	4.0	6.0	0.040
3	6.0	4.0	0.060
4	8.0	2.0	0.080
5	~10	0	0.10

Part B: Calibrate Spectrophotometer

4. Prepare a blank cuvette by filling it ¾ full with deionized water. Handle cuvettes by the ribbed sides, wipe them clean, and ensure no bubbles.
5. Place the blank cuvette in the instrument and follow the instrument's instructions to do the calibration.

VERNIER TECH BOX: SPECTROVIS + LABQUEST 3

1. Connect & Warm Up

- Connect the **SpectroVis Plus** to the USB port of the **LabQuest 3**.
- Launch the LabQuest App. You should see a large "Absorbance" readout.
- Allow the lamp to warm up for 2–3 minutes.

2. Calibration (The "Blank")

- Fill a cuvette 3/4 full with your solvent (Deionized Water) and place it in the spectrometer. **Align the clear sides with the arrow.**
- Tap the red **Absorbance** box on the screen → Select **Calibrate**.
- Tap **Finish Calibration** (or "Skip Warmup" if already warm) → **OK**.

3. Setup for Beer's Law Data (Events with Entry)

- Tap the **Mode: Full Spectrum** box. Change it to **Events with Entry**.
- **Name:** Concentration
- **Units:** M (or mol/L)
- Select **OK**.
- *You are now ready to collect data points. The device will record Absorbance automatically; you must type in the Concentration.*

Part C: Determine Optimal Wavelength (λ_{max})

6. Remove the blank and empty the cuvette.
7. Rinse the cuvette with the 0.10 M standard solution, then fill the cuvette ¾ full with the 0.10 M standard solution.
8. Place the cuvette in the cuvette slot and start data collection.
9. Collect a full spectrum and identify the wavelength of maximum absorbance (λ_{max}). Record this wavelength.
10. Save a copy of the spectrum to include it in your Lab Report.

Part D: Collect Absorbance Data for Standards

11. Set up the spectrometer to collect absorbance data at λ_{max}.
12. Start data collection.
13. For each of the five standard solutions (Test Tubes 1–5):
    a. Rinse the cuvette twice with the solution, then fill it ¾ full.
    b. Wipe the cuvette and place it in the spectrophotometer.
    c. Once the absorbance reading stabilizes, record the absorbance.

DATA SKILL: VISUAL REALITY CHECK

Stop and look at your data values before graphing.

- **The Trend:** As concentration doubles (e.g., from 0.020 M to 0.040 M), your absorbance should roughly double.
- **The Check:** If Tube 3 (0.060 M) has a *lower* absorbance than Tube 2 (0.040 M), you have a physical error (dirty cuvette or mixed-up tubes).
- **The Fix:** Do not include "impossible" points in your regression. Re-measure that specific standard now.

Part E: Analyze Standard Data (Beer's Law Curve)

14. Use the spectrometer software or enter your data into a spreadsheet program to display a graph of absorbance (y-axis) vs. concentration (x-axis).
15. Apply a linear regression curve fit to the data.
    a. The linear regression statistics for these data are for the equation in the form: $y = mx + b$, where x is concentration, y is absorbance, *m* is the slope, and *b* is the y-intercept. One indicator of data quality is the size of *b*: it should be very small if the regression line passes near the origin. The correlation coefficient *r* indicates goodness of fit; a value close to 1.00 means an excellent fit.
    b. The regression line should closely fit the five data points and pass through (or near) the origin of the graph.

DATA SKILL: INTERPRETING REGRESSION STATISTICS

Your regression equation is $y = mx + b$. Here is what the math means physically:

- **The Slope (m):** This represents the sensitivity of your analysis. According to Beer's Law ($A = \epsilon bC$), the slope equals $\epsilon \times b$ (Molar Absorptivity × Path Length).

- **The Intercept (b):** Theoretically, this should be zero (zero concentration = zero light absorbed).
  - If $b > 0.05$: You likely had a fingerprint on the cuvette or your "Blank" wasn't pure water.
  - If $b < -0.05$: You may have calibrated the instrument incorrectly.
- **Correlation (r):** If $r < 0.98$, your pipetting technique was inconsistent.

16. Enter the slope (m), y-intercept (b), and correlation coefficient (r) from the regression statistics in the Data Table. Save the graph to include it in your Lab Report.

Part F: Determine Unknown Concentration

17. Obtain about 5 mL of the unknown $CuSO_4$ solution and record its ID.
18. Rinse the cuvette twice with the unknown solution, fill it ¾ full, wipe it, and place it in the spectrophotometer.
19. Monitor the absorbance until it stabilizes and record this absorbance value in your data table.
20. Use the linear regression results to calculate the concentration of the unknown solution from its absorbance.

DATA SKILL: CALCULATING THE UNKNOWN

Do not estimate the concentration by looking at the graph. You must use your regression equation.

You measured the Absorbance (y) and need to find the Concentration (x). Rearrange the linear equation:

$$y = mx + b \quad \longrightarrow \quad x = \frac{y - b}{m} \tag{1.1.1}$$

Using your specific slope (m) and intercept (b) from Step 16, calculate the precise concentration of your unknown.

21. Record the concentration in the Data Table.

CHEMICAL DISPOSAL

- Dispose of waste as instructed.

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1.2: Beer's Law - Pre-lab

Name: _____ Date: _____

1. A student pipettes 3.00 mL of a 0.150 M stock solution and dilutes it to a final volume of 50.0 mL. What is the final concentration of the diluted solution?

2. What is Beer’s Law, and how does it relate absorbance to concentration in a solution?

3. Why is it important to calibrate the spectrometer with a blank cuvette before measuring the absorbance of the standard solutions?

4. How can the linear regression equation obtained from the Beer’s Law plot be used to determine the concentration of an unknown copper(II) sulfate solution?

5. **Data Logic Check:** A student records the following absorbances for their standards: Without calculating anything, which data point is obviously an "Outlier"? What physical error likely caused the absorbance to drop even though the concentration

increased?

- 0.02 M: 0.150
- 0.04 M: 0.305
- 0.06 M: 0.280 (!!)

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1.3: Beer's Law - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Unknown ID: _____

Wavelength used: _____

Table **1.3.1**: Data Table

Test Tube Number	Concentration [M]	Absorbance
1	0.020	
2	0.040	
3	0.060	
4	0.080	
5	0.10	

Table **1.3.2**: Beer's Law Fit and Unknown Concentration

Parameter	Value
slope, *m* [units =]	
intercept, *b* [units =]	
correlation coefficient, *r* [no units]	
Unknown absorbance [no units]	
Unknown concentration [M]	

Insert an image of the full spectrum of the standard solution

Insert an image of the Beer’s Law Plot

POST-LAB QUESTIONS

1. The procedure directs you to determine the "wavelength of maximum absorbance (λ_{max})" for the copper(II) sulfate solution. Explain why it is important to measure the absorbance of your standard and unknown solutions at this specific wavelength, rather than at an arbitrary wavelength. What advantage does measuring at λ_{max} provide for determining the concentration of the unknown solution?

2. After collecting absorbance-concentration data for the standard solutions, a linear regression curve is fitted to the data, yielding a correlation coefficient (r) and a y-intercept (b). The Introduction states that "A value of 1.00 indicates a nearly perfect fit" for r, and b "is a very small value if the regression line passes through or near the origin." Describe what a low r value (e.g., 0.85) or a significant b value (e.g., 0.10) on your Beer's Law plot might indicate about your experimental data or potential issues during the lab.

Data Analysis Hint

When answering this question, consider the physical reality of the spectrophotometer:

- A low r **value** implies random error (inconsistent pipetting volume).
- A high **y-intercept (b)** implies systematic error (a dirty blank or background interference).

Use these concepts to explain specifically what "went wrong" in the examples provided.

3. The experimental procedure outlines specific instructions for handling cuvettes, such as wiping them with lint-free tissues, handling them only by the ribbed sides, dislodging bubbles, and ensuring light passes through the clear sides. Choose two of these handling instructions and explain how neglecting them would affect your absorbance readings and, consequently, the calculated concentration of your unknown copper(II) sulfate solution. Be specific.

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CHAPTER OVERVIEW

2: Determination of an Equilibrium Constant

PURPOSE

- To apply Le Châtelier's Principle to shift equilibrium conditions and drive the formation of $\mathbf{Fe(SCN)^{2+}}$.
- To use spectrophotometry and Beer's Law to determine the concentration of $\mathbf{Fe(SCN)^{2+}}$ and calculate the proportionality constant k.
- To construct and analyze a calibration curve by plotting absorbance versus concentration and interpret the slope for quantitative analysis.
- To calculate the equilibrium concentrations of $\mathbf{Fe^{3+}}$, $\mathbf{SCN^-}$, and $\mathbf{Fe(SCN)^{2+}}$ in order to determine the equilibrium constant K_c.

INTRODUCTION

This experiment allows you to determine the equilibrium constant for the formation of the iron(III) thiocyanate complex ion, $\mathbf{Fe(SCN)^{2+}}$, and to apply Le Châtelier's Principle to control the position of equilibrium. When $\mathbf{Fe^{3+}}$ and $\mathbf{SCN^-}$ are mixed in aqueous solution, they form the blood-red complex ion $\mathbf{Fe(SCN)^{2+}}$:

$$\mathbf{Fe^{3+}(aq) + SCN^-(aq) \rightleftharpoons Fe(SCN)^{2+}(aq)} \tag{2.1}$$

colorless colorless blood-red

The complex ion is highly colored and easily detected by spectrophotometry. Because the reactants are essentially colorless, absorbance measurements directly correspond to the concentration of $\mathbf{Fe(SCN)^{2+}}$ without interference.

The equilibrium constant for the reaction is given by:

$$K_c = \frac{[\mathbf{Fe(SCN)^{2+}}]_{eq}}{[\mathbf{Fe^{3+}}]_{eq}\,[\mathbf{SCN^-}]_{eq}} \tag{2.2}$$

To determine K_c, the equilibrium concentrations must be known. Beer's Law provides a direct method for determining $\mathbf{Fe(SCN)^{2+}}$ concentration from absorbance measurements:

$$A = kc \tag{2.3}$$

However, the proportionality constant k must first be determined using solutions of *known* $\mathbf{Fe(SCN)^{2+}}$ concentration. This is complicated by the fact that the complex ion participates in the equilibrium itself. Stoichiometric mixing does not guarantee a known concentration of product.

This difficulty is resolved by Le Châtelier's Principle. By adding a large excess of $\mathbf{Fe^{3+}}$, the equilibrium shifts so far to the right that essentially all $\mathbf{SCN^-}$ is converted into $\mathbf{Fe(SCN)^{2+}}$. Under these conditions, the concentration of the complex ion is effectively equal to the initial concentration of the limiting reagent, $\mathbf{SCN^-}$.

In the experiment, you will prepare solutions containing known amounts of $\mathbf{Fe(SCN)^{2+}}$ and measure their absorbances at approximately 450 nm, the wavelength of maximum absorbance. Plotting absorbance versus concentration produces a straight line whose slope is the Beer's Law constant k.

After establishing k, you will measure the absorbance of equilibrium mixtures prepared with various starting concentrations of $\mathbf{Fe^{3+}}$ and $\mathbf{SCN^-}$. From these results, equilibrium concentrations will be determined and used to calculate the equilibrium constant K_c for each trial.

DATA PREP: PRECISION VS. ACCURACY

You will calculate K_c five times. You might expect to get the exact same number every time, but experimental noise prevents this.

The Coefficient of Variation (CV): To determine if your equilibrium constant is truly "constant," we calculate the CV (Standard Deviation divided by the Average). A low CV (below 5%) proves that your technique was precise and the reaction follows the equilibrium model well.

Key Equations

Beer's Law:

$$A = \varepsilon bc \quad \text{or} \quad A = kc \tag{2.4}$$

Dilution Equation:

$$C_1V_1 = C_2V_2 \tag{2.5}$$

2.1: Determination of an Equilibrium Constant - Experiment

2.2: Determination of an Equilibrium Constant - Pre-lab

2.3: Determination of an Equilibrium Constant - Data and Report

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2.1: Determination of an Equilibrium Constant - Experiment

Part A: Determine Beer's Law Constant

1. Obtain five large test tubes. Wash, rinse, and dry the test tubes, then number them 1 through 5.
2. Prepare the five reference solution test tubes listed in the table below. Use a separate pipette (+/- 0.02 mL) to transfer the appropriate volumes of each reagent.
3. Mix each solution thoroughly using a stirring rod. Rinse the stirring rod and dry it between solutions.

Table **2.1.2**: Solution volumes for Beer's Law determination

Test Tube	0.200 M $Fe(NO_3)_3$ [mL]	0.00020 M KSCN [mL]
1	4.0	6.0
2	5.0	5.0
3	6.0	4.0
4	7.0	3.0
5	8.0	2.0

VERNIER TECH BOX: SPECTROMETER SETUP

Read this before proceeding to Step 4.

1. Calibration (The Blank)

- Fill a cuvette 3/4 full with deionized water.
- Tap the red **Absorbance** readout → **Calibrate** → **Finish Calibration**.
- The reading should drop to 0.000 Abs.

2. Determining λ_{max}

- Place **Test Tube 1** (the darkest solution) in the holder.
- Tap **Collect**. When the spectrum appears, tap **Stop**.
- Tap the peak of the graph (should be near 450 nm). Note this wavelength.

3. Measuring Your Samples

- Tap the **Meter Tab** (gauge icon).
- Tap on the "Mode" box and select **Selected Events**. This prevents the number from jumping around while you record it.

4. Use deionized water as the blank to calibrate the spectrophotometer.
5. Obtain a full spectrum of the solution in test tube 1. This should be your most concentrated iron(III) thiocyanate solution.
6. Find the wavelength at the highest point of the peak near 450 nm. Record this value as your λ_{max}. Use this wavelength for all subsequent absorbance measurements.
7. Save your spectrum so that you can include a graph of the absorbance of iron(III) thiocyanate in your report.
8. Measure and record the absorbance of each solution, test tubes 1–5, at your λ_{max}.

Part B: Determine the Equilibrium Constant

1. Wash and dry five test tubes. Number the test tubes 6 through 10.
2. Read the reagent labels carefully before use! Prepare the solutions as described in the table below. Use a separate pipette for each reagent and water and be sure to properly rinse the pipette before each addition.
3. Mix each solution thoroughly using a stirring rod. Rinse the stirring rod and dry it between solutions.

Table **2.1.3**: Solution volumes for Equilibrium Constant determination

Test Tube	0.0020 M $Fe(NO_3)_3$ [mL]	0.0020 M KSCN [mL]	Deionized Water [mL]
6	5.0	1.0	4.0
7	5.0	2.0	3.0
8	5.0	3.0	2.0
9	5.0	4.0	1.0
10	5.0	5.0	0

4. Measure the absorbances at your λ_{max}.

DATA SKILL: REALITY CHECK

Review your data table before leaving the instrument.

- **The Chemical Trend:** In tubes 6 through 10, you are adding progressively more SCN^- reactant. According to Le Châtelier's Principle, this must shift the equilibrium to the right, producing more red complex.
- **The Check:** Your Absorbance values **must increase** from Test Tube 6 to Test Tube 10.
- **The Fix:** If Tube 8 has a lower absorbance than Tube 7, you likely swapped a pipette or mis-mixed the solution. **Remake that solution immediately.**

CHEMICAL DISPOSAL

- All waste containing the thiocyanate ion must be collected in a waste container.

2.1: Determination of an Equilibrium Constant - Experiment is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

2.2: Determination of an Equilibrium Constant - Pre-lab

Name: _____ Date: _____

1. "The equilibrium concentration of $\mathbf{Fe(SCN)^{2+}}$ ions in each reference solution is essentially equal to the concentration of $\mathbf{SCN^-}$ ions in solution before any reaction occurs." Use LeChatelier's Principle to explain why this statement is true.

2. The five reference solutions in Part A are prepared by mixing the 0.200 M $\mathbf{Fe(NO_3)_3}$ solution and the 0.00020 M **KSCN** solution in the amounts listed in the Part A table.

   The concentration of $\mathbf{SCN^-}$ ions in the last reference solution (C_2) before any reaction occurs can be calculated using the so-called "dilution equation," as shown below.

   $$C_1V_1 = C_2V_2 \tag{2.2.1}$$

   C_1 = concentration of solution before mixing = 0.00020 M **KSCN**

   V_1 = volume of solution before mixing = 2.0 mL

   V_2 = final volume of reference solution after mixing = 2.0 + 8.0 mL = 10.0 mL

   Therefore,

   $$C_2 = C_1\left(\frac{V_1}{V_2}\right) = (0.000200\text{ M})\left(\frac{2.0\text{ mL}}{10.0\text{ mL}}\right) = 4.0 \times 10^{-5}\text{ M} \tag{2.2.2}$$

   Use the dilution equation to calculate the concentration of $\mathbf{SCN^-}$ ions in the five reference solutions before any reaction occurs. These will be the same as the $\mathbf{Fe(SCN)^{2+}}$ concentrations in the calibration Test Tubes, 1–5. Enter these values here and copy them into the Part A Data Table as [$\mathbf{Fe(SCN)^{2+}}$].

Table **2.2.1**: Solution volumes and thiocyanate concentrations for Beer's Law

Test Tube	0.200 M Fe^{3+} [mL]	0.00020 M SCN^- [mL]	[SCN^-] after dilution
1	4.0	6.0	
2	5.0	5.0	
3	6.0	4.0	
4	7.0	3.0	
5	8.0	2.0	

Show a sample calculation here:

3. The Part B table summarizes the volumes of Fe^{3+} and SCN^- stock solutions that will be mixed to prepare the test solutions in Part B. Use the dilution equation to calculate the concentrations of ions in each test solution before any reaction occurs, $[Fe^{3+}]_0$ and $[SCN^-]_0$. Enter the results of these calculations in scientific notation here and copy them into the Part B Data Table. *Hint*: The final volume (V_2) of each test solution is 10.0 mL.

Table **2.2.2**: iron(III) and thiocyanate concentrations after dilution, before equilibration

Test Tube	$[Fe^{3+}]_0$	$[SCN^-]_0$
6		
7		
8		
9		
10		

Show a sample calculation here:

4. **Trend Prediction:** In Part B, Test Tube 10 has *five times* as much SCN^- added as Test Tube 6.
   - According to Le Châtelier's Principle, should the solution in Tube 10 be darker or lighter red than Tube 6?
   - If you measure the Absorbance of Tube 10 and it is *lower* than Tube 6, what does that tell you about your solution preparation?

2.2: Determination of an Equilibrium Constant - Pre-lab is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

2.3: Determination of an Equilibrium Constant - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Part A. Determining Beer's Law constant

Table **2.3.1**: Standards concentrations and absorbances

Test Tube	$[Fe(SCN)^{2+}]$	Absorbance (at λ_{max} = _____ nm)
1		
2		
3		
4		
5		

Part B. Determining the Equilibrium Constant

Table **2.3.2**: Ion concentrations (after mixing, before equilibration) and absorbances

Test Tube	$[Fe^{3+}]_0$	$[SCN^-]_0$	Absorbance
6			
7			
8			
9			
10			

Note: These concentrations refer to the concentrations of ions in solution after mixing, before any reaction has occurred. See the PRE-LAB QUESTIONS for calculations.

POST-LAB QUESTIONS

Show a sample calculation for each part when possible.

1. Plot absorbance (*A*) vs. concentration of $Fe(SCN)^{2+}$ for test tubes 1–5. Determine the slope of the line, which is *k* (the Beer's law constant). Paste a picture of your plot here:

2. For test tubes 6–10, calculate the following and enter them into the Results Table below:
   a. concentration of iron(III) thiocyanate at equilibrium [from Beer's law]
   b. concentration of iron(III) at equilibrium: subtract the equilibrium concentration of $Fe(SCN)^{2+}$ ions from the initial concentration of Fe^{3+} ions

> **DATA SKILL: LOGICAL BOUNDARIES**
>
> When calculating equilibrium concentrations, apply these "Sanity Checks" to avoid calculation errors:
>
> - **The Ceiling:** $[Fe^{3+}]_{eq}$ **must be less than** $[Fe^{3+}]_{initial}$. You cannot end up with more iron than you started with.
> - **The Floor:** The concentration cannot be negative. If you get a negative number, check your Beer's Law calculation in Step 2a—you may have calculated a $[FeSCN^{2+}]$ that is physically impossible given your starting amounts.

   c. concentration of thiocyanate at equilibrium: subtract the equilibrium concentration of $Fe(SCN)^{2+}$ ions from the initial concentration of SCN^- ions
   d. K_c

Sample Calculations for Test Tube Number 6:

2. Calculate the average and standard deviation for K_c and enter them in the Results Table

Results Table

Table **2.3.3**: Equilibrium ion concentrations and equilibrium constants

Test Tube	$[Fe(SCN)^{2+}]_{eq}$	$[Fe^{3+}]_{eq}$	$[SCN^{-}]_{eq}$	K_c
6				
7				
8				
9				
10				
			Average of the K_c values:	
			Standard Deviation of the K_c values:	

4. Does your standard deviation indicate that the equilibrium constant is indeed “constant” for this reaction? Explain.

DATA SKILL: IS IT ACTUALLY CONSTANT?

In science, "constant" doesn't mean "identical." There is always noise. To decide if your K_c is truly constant, calculate the **Coefficient of Variation (CV)**:

$$CV = \left(\frac{\text{Standard Deviation}}{\text{Average}}\right) \times 100\% \tag{2.3.1}$$

Interpretation Guide:

- **CV < 5%:** Excellent precision. The constant is definitely constant.
- **CV 5% - 15%:** Acceptable variance for a general chemistry lab.
- **CV > 20%:** The value is drifting. This suggests experimental error or that the reaction is not following the simple equilibrium model perfectly.

In your answer below, cite your CV percentage to support your conclusion.

5. Describe possible sources of error in this experiment and their likely effect on the results.

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CHAPTER OVERVIEW

3: Acid Dissociation Constant

PURPOSE

- To titrate an unknown weak acid using a pH probe.
- To use the titration curve to determine the pK_a (and therefore the K_a) of the acid and determine a likely identity of the acid.

INTRODUCTION

This lab experiment focuses on titrating an unknown weak acid using a pH probe. The primary goal is to determine the acid dissociation constant (K_a) of the unknown weak acid.

The procedure involves titration of the acid using a 0.1 M **NaOH** solution as the titrant, a pH sensor, and a magnetic stirrer to ensure mixing.

Data collection begins by slowly adding the **NaOH** solution to a known volume of the unknown weak acid, while the pH is continuously monitored and recorded. The titration continues until the pH is over 11.

The K_a value is calculated using the Henderson-Hasselbalch equation. Specifically, at the half-equivalence point of the titration, the pH is equal to the pK_a of the weak acid. From the pK_a, the K_a can then be determined. The experiment also involves analyzing the titration graphs to estimate the concentration of the unknown acid from the equivalence point.

Calculating the K_a

For a monoprotic weak acid, **HA**, the Henderson-Hasselbalch equation is

$$\mathrm{pH} = \mathrm{p}K_\mathrm{a} + \log\left(\frac{[\mathrm{A}^-]}{[\mathrm{HA}]}\right) \tag{3.1}$$

where $[\mathrm{HA}]$ and $[\mathrm{A}^-]$ are the equilibrium molar concentrations of the weak acid and its conjugate base, respectively.

At the half-equivalence point, $V_{1/2} = \frac{1}{2}V_{\mathrm{eq}}$,

$$[\mathrm{HA}] = [\mathrm{A}^-], \quad \text{so} \quad \frac{[\mathrm{A}^-]}{[\mathrm{HA}]} = 1, \quad \text{and} \quad \log\left(\frac{[\mathrm{A}^-]}{[\mathrm{HA}]}\right) = 0 \tag{3.2}$$

Therefore, at the half-equivalence point,

$$\mathrm{pH} = \mathrm{p}K_\mathrm{a} \tag{3.3}$$

The K_a can be calculated using the definition of $\mathrm{p}K_\mathrm{a}$,

$$\mathrm{p}K_\mathrm{a} = -\log K_\mathrm{a}, \quad \text{so} \quad K_\mathrm{a} = 10^{-\mathrm{p}K_\mathrm{a}} \tag{3.4}$$

DATA PREP: READING THE INFLECTION POINT

You will determine the equivalence point (V_{eq}) visually from the steep part of your graph.

- **Steepness Matters:** The sharper the jump in pH, the more precisely you can read V_{eq}.
- **Data Density:** If you add large volumes of base (e.g., 1.0 mL) near the equivalence point, you will miss the turn. You need more data points (small drops) in the steep region to get a precise reading.

Key Equations

Henderson-Hasselbalch Equation:

$$\mathrm{pH} = \mathrm{p}K_\mathrm{a} + \log\left(\frac{[\mathrm{A}^-]}{[\mathrm{HA}]}\right) \tag{3.5}$$

Definition of pK_a:

$$\mathrm{p}K_\mathrm{a} = -\log K_\mathrm{a} \quad \text{or} \quad K_\mathrm{a} = 10^{-\mathrm{p}K_\mathrm{a}} \tag{3.6}$$

3.1: Acid Dissociation Constant - Experiment

3.2: Acid Dissociation Constant - Pre-lab

3.3: Acid Dissociation Constant - Data and Report

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3.1: Acid Dissociation Constant - Experiment

SAFETY PRECAUTIONS

- The unknown acid AND sodium hydroxide solutions are caustic. Avoid spilling it on your skin or clothing.

EQUIPMENT AND CHEMICALS NEEDED

Table **3.1.1**: Materials List

Equipment	Chemicals
pH probe	Unknown monoprotic weak acid
buret, buret clamp, and stand	0.1 M $\mathrm{NaOH(aq)}$
magnetic stir bar and stirrer	
150-mL beaker (or smaller)	
250-mL beaker	

EXPERIMENTAL PROCEDURE

Part A: Set Up the Titration

1. Add 30 mL of deionized water to a 100- or 150-mL beaker. Use a pipette bulb (or pipette pump) to pipette 10.00 mL of the unknown weak acid into the beaker with the deionized water.
2. Place the magnetic stirrer on the base of the ring stand. Carefully add the stir bar to the beaker and turn on the stirrer to mix the solution well.
3. Rinse the pH Sensor with deionized water put the sensor into the acid solution. Be careful to avoid striking the fragile tip of the pH probe with the spinning stirrer. You can use a utility clamp or other device to hold the pH probe above the stirrer.

VERNIER TECH BOX: pH TITRATION SETUP

1. Sensor Connections & pH Calibration (All Users)

- Connect the **pH Sensor** to **CH 1**.
- Tap the red **pH** readout → **Calibrate** → **Calibrate Now**.
- **Point 1:** Rinse probe, place in **pH 4.00 Buffer**. Enter "4.00" and tap **Keep**.
- **Point 2:** Rinse probe, place in **pH 7.00 Buffer**. Enter "7.00" and tap **Keep**.
- Tap **OK**.

---

2. Method A: Manual Titration (No Drop Counter)

- Tap **Mode**. Select **Events with Entry**.
- **Name:** Volume. **Units:** mL.
- In this mode, add a small amount of base, wait for pH to stabilize, tap **Keep**, and manually type in the total buret volume reading.

---

3. Method B: Automated Titration (With Drop Counter)

- Connect the **Drop Counter** to **DIG 1**.
- Tap **Mode**. Select **Drop Counting**.
- **CRITICAL STEP: Calibrate Drops**
  - Tap the red **Drop Counting** box → **Calibrate** → **Calibrate Now**.
  - Place a 10-mL graduated cylinder under the drop tip.
  - Open the valve fully. Allow drops to flow until you have collected between 9 and 10 mL. **Close the valve.**

- Enter the **exact volume** of liquid in the cylinder (e.g., 9.40 mL) and tap **Keep**.
- You are now ready. The device knows exactly how big your specific drops are.

4. Obtain approximately 100 mL of 0.1 M **NaOH** solution in a 250 mL beaker. Record the **NaOH** concentration in your data table.
5. Fill the 50-mL buret with the **NaOH** solution and clamp the buret above the beaker.

Part B: Titrate the Unknown

6. You are now ready to begin collecting data. First, check to see that the pH value is acidic. Record the initial buret reading and the pH.
7. Add **NaOH** slowly, no more than 0.50 mL, and record the buret reading and the pH.
8. Continue to add **NaOH** slowly and record buret readings and pH until the pH reaches at least 11. **NOTE**: When you near the equivalence point, you should slow down your additions to only about 0.1–0.2 mL each time. You will know when you are near the equivalence point when the pH starts changing very rapidly.

Part C: Repeat the Titration

9. Repeat the titration with a second sample of 10.00 mL of the unknown acid.
10. Rinse the pH sensor with deionized water before returning it to the storage bottle.

CHEMICAL DISPOSAL

- Dispose of the beaker contents as directed by your instructor.
- Dispose of any waste as directed by your instructor.

3.1: Acid Dissociation Constant - Experiment is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

3.2: Acid Dissociation Constant - Pre-lab

Name: _____ Date: _____

1. Write the acid dissociation reaction for a monoprotic acid, **HA**.

2. Write the expression for K_a.

3. What two species have equal concentrations at the half-equivalence point?

4. A student plans to titrate 20.00 mL of a 0.100 M weak acid solution with 0.100 M **NaOH**. What volume of **NaOH** will be required to reach the **half-equivalence point**?

5. **Data Strategy:** Why does the procedure instruct you to "slow down your additions to only about 0.1-0.2 mL" near the equivalence point?
   Hint: Draw two titration curves—one with 5 data points in the steep region and one with only 1 data point. Which one allows you to determine V_{eq} more precisely?

3.2: Acid Dissociation Constant - Pre-lab is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

3.3: Acid Dissociation Constant - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Table 3.3.1: Titration Data Table

First Titration		Second Titration	
Buret reading (mL)	**pH**	**Buret reading (mL)**	**pH**

POST-LAB QUESTIONS

1. For each titration, create a titration curve using a spreadsheet or data analysis software. Paste your graphs here and comment on the overall appearance of the graphs. What major features are there? How repeatable was the experiment?

2. Estimate the equivalence points, V_{eq}, from the titration curves. From these, calculate the half-equivalence points, $V_{\frac{1}{2}} = \frac{1}{2}V_{eq}$, estimate the pH values (=pK_a) at the half-equivalence points, and calculate the K_a for each curve. Enter all of these values into the Results Table below.

Table **3.3.1**: Results Table

	Trial 1	**Trial 2**
Volume at the equivalence point, V_{eq} [mL]		
Volume at the half-equivalence point, $V_{1/2}$ [mL]		
pK_a (= pH at $V_{1/2}$)		
K_a		

3. Calculate the concentration of the unknown acid from the equivalence volumes. You will get two values: one from the equivalence point on each graph. Report the individual values and the average, and comment on the precision.

DATA SKILL: QUANTIFYING PRECISION (N=2)

Since you only have two trials, you cannot calculate a Standard Deviation. Instead, use **Percent Difference** to quantify your precision:

$$\% \text{ Diff} = \frac{|\text{Trial 1} - \text{Trial 2}|}{\text{Average}} \times 100\% \tag{3.3.1}$$

- **Analysis:** If your % Difference is > 5%, your titration technique was inconsistent. Discuss specifically whether the error came from reading the buret or estimating the equivalence point on the graph.

4. From the estimate of the value of pK_a and K_a from each of the graphs, use a table of acid ionization constants to propose a candidate for the identity of the unknown acid.

DATA SKILL: QUANTIFYING ACCURACY

Once you identify your acid candidate, look up its accepted literature K_a value.

Calculate the **Percent Error** of your experimental average K_a:

$$\% \text{ Error} = \frac{|\text{Experimental} - \text{Literature}|}{\text{Literature}} \times 100\% \tag{3.3.2}$$

Note: It is common for K_a errors to be large (10-20%) because the pKa scale is logarithmic. A small pH error creates a large K_a error.

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CHAPTER OVERVIEW

4: Titration of a Diprotic Acid

PRE-LAB PREPARATION

*Before taking your **Lab Entry Ticket**, ensure you have mastered the following.*

1. Prerequisite Math & Theory

- **Polyprotic Stoichiometry:** A diprotic acid ($\mathbf{H_2A}$) releases two protons.
  - **First Equivalence Point (V_{eq1}):** All $\mathbf{H_2A}$ is converted to $\mathbf{HA^-}$.
  - **Second Equivalence Point (V_{eq2}):** All $\mathbf{HA^-}$ is converted to $\mathbf{A^{2-}}$.
  - **The Rule:** It takes the same amount of base to remove the second proton as the first. Therefore, $V_{eq2} \approx 2 \times V_{eq1}$.
- **Finding pK_a:**
  - pK_{a1} is the pH at $0.5 \times V_{eq1}$.
  - pK_{a2} is the pH at $V_{eq1} + 0.5 \times (V_{eq2} - V_{eq1})$, which simplifies to $1.5 \times V_{eq1}$.

2. Required Technical Skills (See Appendix 13.1)

- **Section B:** Calibrating a pH Probe.
- **Section C (or D):** Titration Data Collection (Drop Counter or Manual).
- **Section I.2:** Using the **Tangent Tool**. Because the first "bump" in a diprotic curve is often small, visually guessing the center is inaccurate. The Tangent tool finds the exact volume where the slope is maximum.

3. Critical Safety

- **Chemical Handling:** Sodium Hydroxide is corrosive. Wear goggles at all times.

PURPOSE

- To titrate an unknown diprotic weak acid using a pH probe.
- To use the titration curve to determine the pK_{a1} and pK_{a2} of the acid and determine a likely identity of the acid.

INTRODUCTION

This lab experiment is designed to investigate the titration of a diprotic weak acid, referred to as $\mathbf{H_2A}$. Unlike monoprotic acids, a diprotic acid is characterized by the presence of two acidic protons that can dissociate in a stepwise manner, each associated with its own acid dissociation reaction and a specific acid dissociation constant, K_{a1} and K_{a2}. The primary goal of this experiment is to explore these dissociations and determine the quantitative values that define the acid's strength and behavior in an aqueous solution.

To achieve this, a potentiometric titration will be performed while continuously monitoring the pH of the acid solution. A strong base, specifically a 0.1 M $\mathbf{NaOH}$ solution, will serve as the titrant, delivered from a buret.

Upon completion of the titration, the collected titrant volume and pH data will be used to generate a titration curve, a graphical representation essential for identifying critical points during the reaction. From this curve, the first and second equivalence points will be determined, which correspond to the complete neutralization of each acidic proton. Subsequently, the pK_{a1} and pK_{a2} values for the diprotic acid will be estimated from the half-equivalence points. These values will then allow for the calculation of K_{a1} and K_{a2}. Ultimately, the experiment aims to estimate the concentration of the unknown diprotic acid and, by comparing the determined pK_a values with a table of known acid ionization constants, propose a plausible identity for the unknown acid.

DATA PREP: THE "DOUBLE VOLUME" RULE

Diprotic acids have two equivalence points. Because the stoichiometry is 1:1 for each proton removal, there is a built-in mathematical check:

- The volume to reach the **second** equivalence point (V_{eq2}) should be exactly **twice** the volume of the **first** equivalence point (V_{eq1}).
- **Data Check:** If your V_{eq2} is not roughly $2 \times V_{eq1}$, you likely misidentified one of the endpoints on your graph.

Key Equations

Henderson-Hasselbalch Equation:

$$\mathrm{pH} = \mathrm{p}K_\mathrm{a} + \log\left(\frac{[\mathrm{A}^-]}{[\mathrm{HA}]}\right) \tag{4.1}$$

Definition of p*K*_a:

$$\mathrm{p}K_\mathrm{a} = -\log K_\mathrm{a} \quad \text{or} \quad K_\mathrm{a} = 10^{-\mathrm{p}K_\mathrm{a}} \tag{4.2}$$

4.1: Titration of a Diprotic Acid - Experiment
4.2: Titration of a Diprotic Acid - Pre-lab
4.3: Titration of a Diprotic Acid - Data and Report

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4.1: Titration of a Diprotic Acid - Experiment

SAFETY PRECAUTIONS

- The unknown acid AND sodium hydroxide solutions are caustic. Avoid spilling it on your skin or clothing.

EQUIPMENT AND CHEMICALS NEEDED

Table **4.1.1**: Materials List

Equipment	Chemicals
pH probe	Unknown diprotic weak acid (H_2A)
buret, buret clamp, and stand	0.1 M $NaOH(aq)$
magnetic stir bar and stirrer	
150-mL beaker (or smaller)	
250-mL beaker	

EXPERIMENTAL PROCEDURE

Part A: Set Up the Titration

1. Add 30 mL of deionized water to a 100- or 150-mL beaker. Use a pipette bulb (or pipette pump) to pipette 20.00 mL of the unknown acid into the beaker with the deionized water.
2. Place the magnetic stirrer on the base of the ring stand. Carefully add the stir bar to the beaker and turn on the stirrer to mix the solution well.
3. Rinse the pH Sensor with deionized water put the sensor into the acid solution. Be careful to avoid striking the fragile tip of the pH probe with the spinning stirrer. You can use a utility clamp or other device to hold the pH probe above the stirrer.

VERNIER TECH BOX: DIPROTIC TITRATION SETUP

1. Sensor Connections & pH Calibration

- Connect the **pH Sensor** to **CH 1**.
- Tap the red **pH** readout → **Calibrate** → **Calibrate Now**.
- **Point 1:** Use **pH 4.00 Buffer**. Enter "4.00" and tap **Keep**.
- **Point 2:** Use **pH 7.00 Buffer**. Enter "7.00" and tap **Keep**.
- Tap **OK**.

---

2. Method A: Manual Titration (No Drop Counter)

- Tap **Mode**. Select **Events with Entry**.
- **Name:** Volume. **Units:** mL.
- In this mode, you manually enter the total buret volume after each increment.
- *Tip: Take smaller steps (0.1 mL) around the expected first equivalence point to capture the curve shape.*

---

3. Method B: Automated Titration (With Drop Counter)

- Connect the **Drop Counter** to **DIG 1**.
- Tap **Mode**. Select **Drop Counting**.
- **Calibration Check:** If you haven't calibrated the drops today, do it now (See Appendix 13.1, Section C).
- Open the valve to a slow drip (1 drop/sec). The system will draw the double-hump curve automatically.

4. Obtain approximately 100 mL of 0.1 M **NaOH** solution in a 250 mL beaker. Record the **NaOH** concentration in your data table.

5. Fill the 50-mL buret with the **NaOH** solution and clamp the buret above the beaker.

Part B: Titrate the Unknown

6. You are now ready to begin collecting data. First, check to see that the pH value is acidic. Record the initial buret reading and the pH.
7. Add **NaOH** slowly, no more than 0.50 mL, and record the buret reading and the pH.
8. Continue to add **NaOH** slowly and record buret readings and pH until the pH reaches at least 11. **NOTE**: When you near the equivalence points, you should slow down your additions to only about 0.1–0.2 mL each time. You will know when you are near an equivalence point when the pH starts changing very rapidly.

DATA SKILL: RESOLUTION CHECK

Before you dump your beaker: Look at your titration curve on the screen (or sketch it).

- **Can you see two distinct steps?**
- Sometimes the first equivalence point is subtle (a small wiggle) rather than a steep cliff. If you missed it by adding base too fast, you will not be able to determine pK_{a1}.
- **Decision:** If your first step is invisible, you must repeat the titration, adding **smaller drops** (0.1 mL) in the early region of the titration.

Part C: Repeat the Titration

9. Repeat the titration with a second sample of 20.00 mL of the unknown acid.
10. Rinse the pH sensor with deionized water before returning it to the storage bottle.

CHEMICAL DISPOSAL

- Dispose of the beaker contents as directed by your instructor.
- Dispose of any waste as directed by your instructor.

4.1: Titration of a Diprotic Acid - Experiment is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

4.2: Titration of a Diprotic Acid - Pre-lab

Name: _____ Date: _____

1. Write the two acid dissociation reactions for a diprotic acid, H_2A.

2. Write the expressions for K_{a1} and K_{a2}.

3. What two species have equal concentrations at the first half-equivalence point and what two species have equal concentrations at the second half-equivalence point?

4. **Prediction:** A student titrates 20.00 mL of 0.100 M $\mathbf{H_2A}$ with 0.100 M **NaOH**.
   - Calculate the volume of **NaOH** needed to reach the **first** equivalence point.
   - Calculate the volume needed to reach the **second** equivalence point.

- Sketch the expected shape of the curve, labeling the volumes at 10 mL, 20 mL, and 40 mL.

4.2: Titration of a Diprotic Acid - Pre-lab is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

4.3: Titration of a Diprotic Acid - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Table 4.3.1: Titration Data Table

https://chem.libretexts.org/@go/page/544322

First Titration		Second Titration	
Buret reading (mL)	pH	Buret reading (mL)	pH

POST-LAB QUESTIONS

1. For each titration, create a titration curve using a spreadsheet or data analysis software. Insert your graphs here and comment on the overall appearance of the graphs. What major features are there? How repeatable was the experiment?

2. Estimate the equivalence points, $V_{eq}^{(1)}$ and $V_{eq}^{(2)}$, from the titration curves. These are most easily estimated by looking for the largest slope of the titration curve. From these, calculate the half-equivalence points, $V_{1/2}^{(1)}$ and $V_{1/2}^{(2)}$, estimate the pH values (=pK_{a1} and pK_{a2}) at the half-equivalence points, and calculate K_{a1} and K_{a2} for each curve. Enter all of these values into the Results Table below.

Table 4.3.2: Results Table

	Calculation	Trial 1	Trial 2
Volume at the first equivalence point [mL]	$V_{eq}^{(1)}$		
Volume at the first half-equivalence point [mL]	$V_{1/2}^{(1)} = \frac{1}{2} V_{eq}^{(1)}$		

3. Calculate the concentration of the unknown acid from the equivalence volumes. You will get four values: one from the first and second equivalence points on each graph. Report the individual values and the average, and comment on the precision. Show your calculations.

DATA SKILL: INTERNAL CONSISTENCY

You calculated the concentration of acid four different ways (using V_{eq1} and V_{eq2} for two trials). Theoretically, all four should be identical.

- **The Check:** Calculate the **Percent Difference** between the concentration derived from V_{eq1} and the concentration derived from V_{eq2}.
- **Analysis:** If the difference is > 5%, which equivalence point do you trust more? (Usually V_{eq2} is steeper and easier to read). Justify your choice in your conclusion.

4. From the estimate of the value of pK_{a1} and pK_{a2} from each of the graphs, use a table of acid ionization constants to propose a candidate for the identity of the unknown acid.

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CHAPTER OVERVIEW

5: Investigating Buffers

PRE-LAB PREPARATION

*Before taking your **Lab Entry Ticket**, ensure you have mastered the following.*

1. Prerequisite Math & Theory

- **Henderson-Hasselbalch Equation:** You must be able to solve for the mass of Sodium Acetate needed to achieve a specific pH.

$$\mathrm{pH} = \mathrm{p}K_a + \log\left(\frac{[\mathrm{Base}]}{[\mathrm{Acid}]}\right) \tag{5.1}$$

  - *Rearrangement Skill:* If you know the target pH and the pK_a, can you solve for the ratio $[\mathrm{Base}]/[\mathrm{Acid}]$?
  - *Mass Calculation:* Once you have the target Molarity of the base (Acetate), use $M \times V \times \text{Molar Mass}$ to find grams.
- **Buffer Capacity:** Understand that a buffer made with **1.0 M** components resists pH change 10x better than a buffer made with **0.1 M** components, even if they have the same initial pH.

2. Required Technical Skills (See Appendix 13.1)

- **Section B:** Calibrating a pH Probe.
- **Section D (or C):** Manual Titration (Events with Entry) is useful here to record "Volume Added" vs "pH," even though we aren't finding an equivalence point.

3. Critical Safety

- **Acids and Bases:** You are working with both concentrated Acetic Acid and Strong Base (NaOH). Goggles are non-negotiable.

PURPOSE

- To use the Henderson-Hasselbalch equation to determine the amounts of acetic acid and sodium acetate needed to prepare two acidic buffer solutions (Buffer A and Buffer B), and then prepare these solutions.
- To determine the buffer capacities of the prepared Buffer A and Buffer B by adding solutions of NaOH (strong base) and HCl (strong acid).

INTRODUCTION

A buffer is a mixture of a weak acid and its conjugate base, or a weak base and its conjugate acid. A buffer's function is to absorb acids ($\mathrm{H^+}$ or $\mathrm{H_3O^+}$ ions) or bases ($\mathrm{OH^-}$ ions) so that the system's pH changes very little.

Buffers are critical in many systems. Blood plasma, a natural example in humans, serves as a bicarbonate buffer, maintaining the pH of blood between 7.2 and 7.6.

By design, a buffer is a system at equilibrium. For example, a buffer can be prepared with nitrous acid, $\mathrm{HNO_2}$. The weak acid establishes an aqueous equilibrium as shown below.

$$\mathrm{HNO_2\,(aq) + H_2O\,(l) \rightleftharpoons H_3O^+\,(aq) + NO_2^-\,(aq)} \tag{5.2}$$

The equilibrium constant expression is shown below.

$$K_a = \frac{[\mathrm{H_3O^+}]\,[\mathrm{NO_2^-}]}{[\mathrm{HNO_2}]} \tag{5.3}$$

To prepare a buffer system with nitrous acid, a salt containing the conjugate base is added, such as sodium nitrite ($\mathrm{NaNO_2}$). The resulting system is a mixture of $\mathrm{HNO_2}$ and $\mathrm{NO_2^-}$ ions. The nitrous acid will neutralize additional hydroxide ions, and the nitrite ion

will neutralize additional hydronium ions.

A variation of the equilibrium expression above, called the Henderson-Hasselbalch equation, is the best reference in preparing a buffer solution. For our nitrous acid–nitrite buffer example, the Henderson-Hasselbalch equation is shown below.

$$\mathrm{pH} = \mathrm{p}K_a + \log\left(\frac{[\mathrm{NO_2^-}]}{[\mathrm{HNO_2}]}\right) \tag{5.4}$$

The pH range in which a buffer solution is effective is generally considered to be ±1 of the pK_a.

In this experiment, you will use the Henderson-Hasselbalch equation to determine the amount of acetic acid and sodium acetate needed to prepare two acidic buffer solutions. You will then prepare the buffers and test their buffer capacities by adding solutions of **NaOH** and **HCl**.

DATA PREP: THE LOGARITHMIC RULE

The Henderson-Hasselbalch equation tells us that pH depends on the **logarithm of the ratio** of base to acid.

- **The "Sweet Spot":** When $[\mathrm{Base}] = [\mathrm{Acid}]$, the ratio is 1. Since $\log(1) = 0$, the pH equals the pK_a. This is the most effective region of the buffer.
- **The "danger zone":** If the ratio drops below 0.1 or goes above 10, the log term shifts by ± 1 unit, and the buffer capacity collapses.

Key Equations

Henderson-Hasselbalch Equation:

$$\mathrm{pH} = \mathrm{p}K_a + \log\left(\frac{[\mathrm{A^-}]}{[\mathrm{HA}]}\right) \tag{5.5}$$

5.1: Investigating Buffers - Experiment

5.2: Investigating Buffers - Pre-lab

5.3: Investigating Buffers - Data and Report

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5.1: Investigating Buffers - Experiment

SAFETY PRECAUTIONS

Sodium hydroxide and hydrochloric acid are both caustic. Handle them with care.

EQUIPMENT AND CHEMICALS NEEDED

Table 5.1.1: Materials List

Equipment	Chemicals
50-mL burets (2)	sodium acetate, $NaC_2H_3O_2$
buret clamp and ring stand	0.1 M acetic acid, $HC_2H_3O_2$
pH probe with electrode support and microstirrer	1.0 M acetic acid, $HC_2H_3O_2$
LabQuest	0.5 M sodium hydroxide, $NaOH$
graduated cylinder	0.5 M hydrochloric acid, HCl
150-mL beaker	
magnetic stirrer	

EXPERIMENTAL PROCEDURE

Part A: Prepare and Test Buffer Solution A

1. Use your calculations from the Pre-Lab Exercise to prepare 50 mL of Buffer A. Weigh out the precise mass of sodium acetate and dissolve it in 50.0 mL of 0.1 M acetic acid solution.
2. Set up two burets, buret clamp, and ring stand. Rinse and fill one buret with 0.5 M $NaOH$ solution. Rinse and fill the second buret with 0.5 M HCl solution.
3. Use a graduated cylinder to measure 10.0 mL of the Buffer A solution into a 150 mL beaker and add 15 mL of distilled water. Place the beaker on a magnetic stirrer, beneath the buret of $NaOH$, and add a stirring bar.
4. Prepare the pH Sensor and use an electrode support or utility clamp to suspend the Sensor in the pH 4 buffer solution. Make sure the stirring bar does not strike the pH sensor.

VERNIER TECH BOX: BUFFER TESTING

1. Calibration Check

- Before starting, dip your probe in pH 4.00 buffer. If it reads between 3.90 and 4.10, you are good. If not, Calibrate (**Appendix 13.1, Section B**).

2. Data Collection Mode

- **Option A (Simple Meter):** Just read the live pH value on the screen and write it in your notebook. You do not need to hit "Collect."
- **Option B (Digital Table):** To create a nice data table automatically:
  - Set Mode to **Events with Entry**.
  - Name the X-axis "Volume Added" (mL).
  - Tap **Collect**. Tap **Keep** after every increment of NaOH, typing in the total volume (0.5, 1.0, 1.5...).

5. Collect data. You will slowly and carefully add 0.50 M $NaOH$ solution to the sample of Buffer A.
   a. Take an initial pH reading of the buffer solution. Allow the pH readings to stabilize and then record the initial buret reading and the initial pH value in the titration data table.

b. Add a small amount of the **NaOH** solution, up to 0.50 mL. When the pH stabilizes, record the buret reading and the pH value in the titration data table.
c. Continue adding the **NaOH** solution in small increments that raise the pH consistently and enter the buret reading after each increment. Your goal is to raise the pH of the buffer by precisely 2 pH units.
d. When the pH of the buffer solution is precisely 2 units greater than the initial reading, continue to add the **NaOH** solution in small increments until you have reached and passed the equivalence point of the titration.

6. Repeat step 5, using a fresh 10.0 mL sample of the Buffer A solution. For this second trial, titrate the buffer with 0.5 M **HCl** solution. Carefully add **HCl** in small increments until the pH of the solution has been lowered by precisely 2 units. Record the buret readings and pH values in the data table.

DATA SKILL: CAPACITY CHECK

Compare Buffer A and Buffer B immediately.

- You should observe that Buffer B requires significantly more drops of **NaOH** or **HCl** to change the pH than Buffer A did.
- **The Trend:** If Buffer B changed pH just as fast as Buffer A, you likely used the 0.1 M acid for both solutions by mistake. **Check your stock bottles.**

Part B: Prepare and Test Buffer Solution B

1. Use your calculations from the Pre-Lab Exercise to prepare 50 mL of Buffer B. Weigh out the precise mass of sodium acetate and dissolve it in 50.0 mL of 1.0 M acetic acid solution. If necessary, refill the burets of **NaOH** and **HCl** solution.
2. Use a graduated cylinder to measure 10.0 mL of the Buffer B solution and add 15 mL of distilled water. Repeat the necessary steps to test Buffer B like the Part A trials, recording your data in the data table.

CHEMICAL DISPOSAL

- Dispose of the reaction mixture as directed.
- Rinse the pH Sensor with distilled water in preparation for the next trial or storage.

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5.2: Investigating Buffers - Pre-lab

Name: _____ Date: _____

1. Use the Henderson-Hasselbalch equation to perform the following calculations. The K_a of acetic acid is 1.8×10^{-5}. Review your calculations with your instructor before preparing the buffer solutions.
   a. Buffer A: Calculate the mass of solid sodium acetate required to mix with 50.0 mL of 0.1 M acetic acid to prepare a pH 4 buffer. Show your work here and record the mass in your data table.

   b. Buffer B: Calculate the mass of solid sodium acetate required to mix with 50.0 mL of 1.0 M acetic acid to prepare a pH 4 buffer. Show your work here and record the mass in your data table.

2. A chemist wishes to make a buffer with a pH of 6.10. Using a table of common weak acids, which acid/conjugate base pair would be the best choice? Explain.

3. **Capacity Prediction:** Buffer B has 1.0 M acetic acid, while Buffer A has only 0.1 M acetic acid.
   - If Buffer B is 10 times more concentrated, how many times more **NaOH** should it be able to neutralize compared to Buffer A?
   - Explain why the pH of both buffers is initially the same despite the concentration difference.

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5.3: Investigating Buffers - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Table **5.3.1**: Buffer A Titration Data Table

NaOH titration of Buffer A			HCl titration of Buffer A	
Buret reading (mL)	**pH**		**Buret reading (mL)**	**pH**

1. Use a spreadsheet or other data analysis software to make plots of each of the titrations. Insert your graphs here:
2. By looking at the data and the plots
   a. Determine the volume of titrant required to change the pH of the solutions by 2 units. Enter these values in the Results Table.
   b. Determine the volume of titrant required to reach the end-point for the NaOH titrations.

Table **5.3.2**: Results Table

	Buffer A	Buffer B
Mass of $NaC_2H_3O_2$ used to prepare buffer (g)		
Volume of buffer prepared (mL)	50.0	50.0
Molar concentration of $HC_2H_3O_2$ in buffer (M)	0.1	1.0
Initial pH of buffer		
Volume of 0.5 M $NaOH$ to raise pH by 2 units (mL)		
Volume of 0.5 M $NaOH$ at end-point (mL)		
Volume of 0.5 M HCl to lower pH by 2 units (mL)		

POST-LAB QUESTIONS

1. Write reaction equations to explain how your acetic acid-acetate buffer reacts with an acid and reacts with a base.

2. Buffer capacity has a rather loose definition, yet it is an important property of buffers. A commonly seen definition of buffer capacity is: “The amount of H^+ or OH^- that can be neutralized before the pH changes to a significant degree.”

DATA SKILL: CALCULATING BUFFER CAPACITY (β)

Instead of a "loose definition," calculate the specific **Buffer Capacity (β)** using the formula:

$$\beta = \frac{\text{moles of strong acid/base added}}{(\Delta\text{pH}) \times (\text{Volume of Buffer in L})} \tag{5.3.1}$$

Task: Calculate β for both Buffer A and Buffer B. Is the capacity of Buffer B roughly 10 times that of Buffer A? Discuss the percent difference from the expected 10x ratio.

Use your data to determine the buffer capacity of Buffer A and Buffer B. Express your answer in units of moles of strong acid or moles of strong base.

3. Say, for example, that you had prepared a Buffer C, in which you mixed 8.203 g of sodium acetate, $NaC_2H_3O_2$, with 100.0 mL of 1.0 M acetic acid.
   a. What would be the initial pH of Buffer C?

b. If you add 5.00 mL of 0.5 M **NaOH** solution to 20.0 mL each of Buffer B and Buffer C, which buffer's pH would change less? Explain.

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CHAPTER OVERVIEW

6: Solubility Product

PRE-LAB PREPARATION

*Before taking your **Lab Entry Ticket**, ensure you have mastered the following.*

1. Prerequisite Math & Theory

- **The 1:6 Stoichiometry Rule:** In this titration, the math is not 1:1.
  - Reaction 1: $1\ \text{mol}\ IO_3^-$ produces $3\ \text{mol}\ I_2$.
  - Reaction 2: $3\ \text{mol}\ I_2$ consumes $6\ \text{mol}\ S_2O_3^{2-}$.
  - **Calculation:** $\text{Moles of Iodate} = \frac{\text{Moles of Thiosulfate}}{6}$. If you miss this factor, your K_{sp} will be wrong by a factor of 216!
- **K_{sp} from Solubility:** If the molar solubility is s, then $[Ca^{2+}] = s$ and $[IO_3^-] = 2s$. The expression becomes $K_{sp} = 4s^3$.

2. Required Technical Skills

- **Gravity Filtration:** You must filter the solution through dry paper into a dry beaker. Any extra water will dilute your saturated solution, ruining the result.
- **Starch Endpoint:** The endpoint is the disappearance of the blue-black color, leaving a colorless solution.

3. Critical Safety

- **Strong Acid:** You will add 3 M HCl to the flask. This releases Iodine, which can stain skin. Wear gloves.

PURPOSE

- To prepare a saturated solution of calcium iodate.
- To determine the concentration of a saturated solution of calcium iodate using a redox titration.
- To determine the solubility product for calcium iodate and compare it to the known value.

INTRODUCTION

In chemistry, the solubility of a compound refers to the maximum amount of solute that can dissolve in a given solvent at equilibrium, forming a saturated solution. For slightly soluble ionic compounds, this equilibrium is described by the solubility product constant (K_{sp}), a fundamental equilibrium constant that quantifies the extent to which a solid dissolves in solution.

For calcium iodate, $Ca(IO_3)_2$, the equilibrium in a saturated solution is represented as:

$$Ca(IO_3)_2\,(s) \rightleftharpoons Ca^{2+}\,(aq) + 2\,IO_3^-\,(aq) \tag{6.1}$$

The solubility product expression is given by:

$$K_{sp} = [Ca^{2+}][IO_3^-]^2 \tag{6.2}$$

where $[Ca^{2+}]$ and $[IO_3^-]$ are the molar concentrations of the dissociated ions in equilibrium.

To determine the solubility of calcium iodate experimentally, we need to measure the concentration of iodate ions in a saturated solution. This is achieved through a redox titration using a standardized sodium thiosulfate ($Na_2S_2O_3$) solution in the presence of potassium iodide (KI) and starch indicator.

The reaction between iodate ions and potassium iodide in acidic solution generates molecular iodine:

$$IO_3^- + 5\,I^- + 6\,H^+ \longrightarrow 3\,I_2 + 3\,H_2O \tag{6.3}$$

The molecular iodine (I_2) is then titrated with sodium thiosulfate, which reduces it back to iodide:

$$I_2 + 2\,S_2O_3^{2-} \longrightarrow 2\,I^- + S_4O_6^{2-} \tag{6.4}$$

where $S_4O_6^{2-}$ is the tetrathionate ion.

Since the amount of thiosulfate used directly correlates with the amount of iodate originally present in solution, this titration allows us to calculate the molarity of iodate ions, determine the solubility of calcium iodate, and compute its solubility product constant (K_{sp}).

Additionally, this experiment explores the concept that the solubility of a slightly soluble salt is independent of the volume of water used to dissolve it, meaning that different volumes of solvent should yield the same solubility value when equilibrium is reached.

DATA PREP: THE 6:1 RATIO

In this redox titration, the stoichiometry is not 1:1. You must trace the electrons carefully:

- Step 1: **1 mole** of IO_3^- produces **3 moles** of I_2.
- Step 2: Each mole of I_2 consumes **2 moles** of thiosulfate ($S_2O_3^{2-}$).
- **Result:** It takes **6 moles** of thiosulfate to titrate every **1 mole** of iodate. If you forget this factor of 6, your calculated K_{sp} will be off by a factor of 6^3 (216 times)!

Key Equations

Solubility Product Constant (K_{sp}):

$$K_{sp} = [Ca^{2+}][IO_3^-]^2 \tag{6.5}$$

Titration Stoichiometry (The 1:6 Ratio):

Because 1 mole of iodate produces 3 moles of iodine, which consume 6 moles of thiosulfate:

$$\text{mol } IO_3^- = \frac{1}{6} \times (\text{mol } S_2O_3^{2-}) \tag{6.6}$$

Calculating K_{sp} from Molar Solubility (s):

If s is the molar solubility of $Ca(IO_3)_2$, then $[Ca^{2+}] = s$ and $[IO_3^-] = 2s$.

$$K_{sp} = (s)(2s)^2 = 4s^3 \tag{6.7}$$

6.1: Solubility Product - Experiment
6.2: Solubility Product - Pre-lab
6.3: Solubility Product - Data and Report

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6.1: Solubility Product - Experiment

SAFETY PRECAUTIONS

- Hydrochloric acid (HCl) is corrosive; avoid skin contact and inhalation.

EQUIPMENT AND CHEMICALS NEEDED

Table 6.1.1: Materials List

Equipment	Chemicals
150-mL beaker	calcium nitrate, $Ca(NO_3)_2$
250-mL Erlenmeyer flask	potassium iodate, KIO_3
10-mL graduated cylinder	Standardized 0.025 M sodium thiosulfate, $Na_2S_2O_3$, solution
10-mL pipet	potassium iodide KI
Buret and buret clamp	3 M hydrochloric acid HCl
Filter paper	1 % starch solution
Funnel	

EXPERIMENTAL PROCEDURE

Part A: Prepare Saturated Solutions of Calcium Iodate

1. Dissolve 5 g $Ca(NO_3)_2$ in 20 mL water in a small beaker. Use 50 mL of water to dissolve 2.5 g KIO_3 in a 150-mL beaker.
2. Add the calcium nitrate solution to the potassium iodate solution and stir to form a precipitate.
3. Let the mixture stand, then filter the precipitate using gravity filtration. Rinse with deionized water.
4. Divide the precipitate and put half of it into a small beaker. Keep the other half of the precipitate in case of emergency.
5. Add 60 mL of deionized water to the precipitate in the beaker. This is your saturated solution. Stir well and let stand for 30 minutes.
6. Prepare a buret with the standardized $Na_2S_2O_3$ solution. Be sure to record the exact molarity.

Part B: Analyze the Saturated Solutions

1. Filter the saturated solution into a **dry** beaker using a **dry** funnel and **dry** filter paper.

TECHNIQUE ALERT: THE "DRY" RULE

Why everything must be DRY:

- You are analyzing a **Saturated Solution**. This means it holds the maximum possible amount of dissolved ions.
- If your funnel, filter paper, or receiving beaker has **water droplets** in it, you are adding solvent.
- **The Result:** The solution becomes unsaturated (diluted). Your calculated K_{sp} will be too low.
- **The Fix:** If you washed your beaker, dry it completely with a paper towel before filtering.

2. Rinse a 10-mL pipet with deionized water and then with the filtrate.
3. Pipet 10.0 mL of the filtrate into a 125-mL Erlenmeyer flask.
4. Add 20 mL of deionized water, 1.5 g of KI, and 15 drops of 3 M HCl. Swirl to mix.
5. Add 10 drops of the starch solution to the flask.
6. Record the initial buret reading.
7. Perform a trial titration by adding $Na_2S_2O_3$ in 1-mL increments until the solution becomes colorless.
8. Record the final buret reading.

9. Repeat these steps, but this time perform an exact titration, using the refined volume from the trial titration. Add drop-wise near the endpoint.
10. Repeat the exact titration on another 10 mL sample of the saturated solution.

CHEMICAL DISPOSAL

- Do not pour chemical waste down the sink. Dispose of waste solutions as instructed by your instructor or lab supervisor.
- Used filter paper and precipitate should be discarded in designated solid waste containers.

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6.2: Solubility Product - Pre-lab

Name: _____ Partner(s): _____ Date: _____

1. Describe the role of potassium iodide and hydrochloric acid in the titration process. Why are they necessary?

2. Explain why starch is used as an indicator in this experiment. What color changes indicate the endpoint of the titration?

3. **Experimental Design:** In Part B, you filter the saturated solution into a **dry** beaker.
   - If your beaker was wet with 1 mL of water, would the concentration of IO_3^- you measure be too high, too low, or unaffected?
   - Would this make your calculated K_{sp} too high or too low? Explain.

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6.3: Solubility Product - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Concentration sodium thiosulfate solution (M): _____

Table **6.3.1**: Trial Titration Data

Trial titration start (mL)
Trial titration end (mL)
Trial titration volume used (mL)

Table **6.3.2**: Exact Titration Data and Calculations

	First Exact Titration	Second Exact Titration
Titration start (mL)		
Titration end (mL)		
Titration volume used (mL)		
Moles $Na_2S_2O_3$ used (mol)		
Moles of IO_3^- (mol)		
Concentration of IO_3^- (M)		
Moles of Ca^{2+} (mol)		
Solubility = concentration of Ca^{2+} (M)		

Average Solubility (M): _____

Solubility product constant, K_{sp}: _____

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POST-LAB QUESTIONS

Show a sample calculation for each part

1. Calculate the number of moles and the molarity of iodate ions from each exact titration. Remember the mole relationship between the thiosulfate and iodine, and the mole relationship between iodine and iodate ions. Enter your answers in the table.

> **DATA SKILL: THE 6x CHECK**
>
> **Stop and check your moles.**
>
> - Your moles of thiosulfate ($S_2O_3^{2-}$) should be roughly **6 times larger** than your moles of iodate (IO_3^-).
> - If your answer for moles of IO_3^- is *larger* than your moles of thiosulfate, you divided when you should have multiplied (or vice versa).

2. Calculate the number of moles and the concentration of calcium ions in the saturated solution. This is the solubility of calcium iodate. Enter your answers in the table.

3. You have calculated two solubility values. Use them to distinguish between precision and accuracy.

> **DATA SKILL: PRECISION VS. ACCURACY**
>
> - **Precision (Repeatability):** Calculate the **Percent Difference** between your two solubility values. A difference $< 2\%$ indicates excellent titration technique.
> - **Accuracy (Truth):** In Question 4, you will compare your average to the literature value. It is possible to have high precision (consistent titrations) but low accuracy (systematic error, like a wet beaker).

4. The standard value for the solubility of calcium iodate is 7.90×10^{-3} M. How well does your average solubility value compare? What is your percent error? Comment on your accuracy.

5. Calculate the Solubility Product Constant, K_{sp}, for calcium iodate using your average solubility.

6. Discuss the effects of errors on the solubility of calcium iodate introduced from the following sources if extraneous water is introduced from wet filter paper or wet funnels, or by washing the precipitate during the final filtration of the calcium iodate.

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CHAPTER OVERVIEW

7: Thermodynamics of Borax

PRE-LAB PREPARATION

*Before taking your **Lab Entry Ticket**, ensure you have mastered the following.*

1. Prerequisite Math & Theory

- **The van 't Hoff Equation:** You will graph your data to find enthalpy and entropy.

$$\ln K = -\frac{\Delta H^\circ}{R}\left(\frac{1}{T}\right) + \frac{\Delta S^\circ}{R} \tag{7.1}$$

  - **Slope:** $m = -\Delta H^\circ / R$
  - **Intercept:** $b = \Delta S^\circ / R$
  - **The Trap:** R is in Joules (8.314 J/mol K), but ΔH is usually reported in kJ. Don't forget the factor of 1000.
- **Stoichiometry:** The tetraborate anion ($B_4O_5(OH)_4^{2-}$) accepts **2 protons** from HCl. Therefore:

$$\text{moles Borax} = 0.5 \times \text{moles HCl} \tag{7.2}$$

2. Required Technical Skills (See Appendix 13.1)

- **Section A:** Measuring Temperature (without hitting "Collect").
- **Section D (or C):** Titration Data Collection (Manual or Drop Counter).
- **Data Skill:** Linearizing data in Excel (converting $T \rightarrow 1/T$ and $K_{sp} \rightarrow \ln K$).

3. Critical Safety

- **Thermal Hazard:** You are heating solutions to 60°C. Glassware will be hot.
- **Chemicals:** 0.20 M HCl is an irritant.

PURPOSE

- To determine the thermodynamic properties (ΔG°, ΔH°, and ΔS°) of borax dissolution by measuring its solubility at various temperatures.
- To calculate the solubility product ($\ln K_{sp}$) of borax at different temperatures using acid-base titration.
- To apply the relationship between $\ln K_{sp}$ and $\frac{1}{T}$ to extract enthalpy and entropy changes for a sparingly soluble salt.
- To gain hands-on experience in solution preparation, equilibrium analysis, and titration techniques as tools for thermodynamic investigation.

INTRODUCTION

The free energy change of a chemical process is proportional to the natural logarithm of its equilibrium constant according to the equation:

$$\Delta G^\circ = -RT \ln K \tag{7.3}$$

where R is the ideal gas constant, 8.314 J/mol·K, and T is the absolute temperature in kelvins. The equilibrium constant, K, is expressed for the equilibrium system when the reactants and products are in their standard states. For a sparingly soluble salt in an aqueous solution, the precipitate and the ions in the solution correspond to the standard states of the reactants and products, respectively.

For instance, the standard state equilibrium silver chromate is

$$Ag_2CrO_4(s) \rightleftharpoons 2Ag^+(aq) + CrO_4^{2-}(aq) \tag{7.4}$$

The solubility product, $\ln K$, is set equal to the product of the molar concentrations of the ions, all raised to the power of their respective coefficients in the balanced equation. This is the mass action expression for the system:

$$K_{sp} = [\mathrm{Ag^+}]^2[\mathrm{CrO_4}^{2-}] \tag{7.5}$$

and the free energy change for the equilibrium is

$$\Delta G^\circ = -RT \ln K_{sp} = -RT \ln [\mathrm{Ag^+}]^2[\mathrm{CrO_4}^{2-}] \tag{7.6}$$

The free energy change of a chemical process is also a function of the enthalpy change and the entropy change of the process:

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ \tag{7.7}$$

When the two free energy expressions are set equal, then

$$-RT \ln K_{sp} = \Delta H^\circ - T\Delta S^\circ \tag{7.8}$$

Solving for $\ln K$ and rearranging yields

$$\ln K_{sp} = -\frac{\Delta H^\circ}{R}\left(\frac{1}{T}\right) + \frac{\Delta S^\circ}{R} \tag{7.9}$$
$$y = m\,x + b$$

This equation can prove valuable in determining the thermodynamic properties of a chemical system, such as that of a sparingly soluble salt. A linear relationship exists when the values of $\ln K_{sp}$ obtained at various temperatures are plotted as a function of the reciprocal temperature, $\frac{1}{T}$. The slope of the line equals $-\frac{\Delta H^\circ}{R}$, and the y-intercept equals $\frac{\Delta S^\circ}{R}$. Since R is a constant, the ΔH° and the ΔS° for the equilibrium system can easily be calculated.

The Borax System

Borax is often given the name sodium tetraborate decahydrate and the formula $\mathrm{Na_2B_4O_7 \cdot 10\,H_2O}$. However, according to its chemical behavior, a more defining formula for borax is $\mathrm{Na_2B_4O_5(OH)_4 \cdot 8\,H_2O}$. The anion, tetraborate ($\mathrm{B_4O_7}^{2-}$), has the structure shown in the figure.

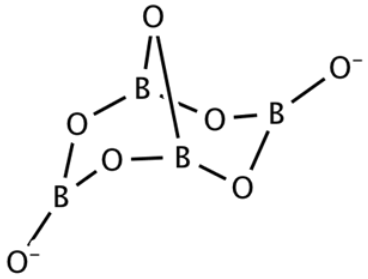


Figure 7.1: The tetraborate anion

Borax dissolves and dissociates in water according to the equation:

$$\mathrm{Na_2B_4O_5(OH)_4 \cdot 8\,H_2O}(s) \rightleftharpoons \mathrm{2\,Na^+}(aq) + \mathrm{B_4O_5(OH)_4^{2-}}(aq) + \mathrm{8\,H_2O}(l) \tag{7.10}$$

The mass action expression, set equal to the solubility product at equilibrium, for the solubility of borax is:

$$K_{sp} = [\mathrm{Na^+}]^2[\mathrm{B_4O_5(OH)_4^{2-}}] \tag{7.11}$$

The tetraborate anion is the conjugate base of the weak acid boric acid and is capable of accepting two protons from a strong acid in an aqueous solution:

$$\mathrm{B_4O_5(OH)_4^{2-}}(aq) + \mathrm{2\,H^+}(aq) + \mathrm{3\,H_2O}(l) \rightleftharpoons \mathrm{4\,H_3BO_3}(aq) \tag{7.12}$$

Therefore, the molar concentration of the tetraborate anion in a saturated borax solution can be measured with titration with a standardized hydrochloric acid solution.

This analysis is also a measure of the molar solubility of borax in water at a given temperature; according to stoichiometry, one mole of the tetraborate anion forms for every mole of borax that dissolves. Also, according to the stoichiometry of the dissolution of the borax (see equation 7.10), the molar concentration of the sodium ion in the saturated solution is twice that of the experimentally determined tetraborate anion concentration:

$$[\mathrm{Na^+}] = 2 \times [\mathrm{B_4O_5(OH)_4^{2-}}] \tag{7.13}$$

The solubility product for borax at a measured temperature is, therefore,

$$\begin{aligned} K_{sp} &= [\mathrm{Na^+}]^2[\mathrm{B_4O_5(OH)_4^{2-}}] \\ K_{sp} &= [2 \times \mathrm{B_4O_5(OH)_4^{2-}}]^2[\mathrm{B_4O_5(OH)_4^{2-}}] \\ K_{sp} &= 4 \times [\mathrm{B_4O_5(OH)_4^{2-}}]^3 \\ K_{sp} &= 4 \times (\text{the molar solubility of borax})^3 \\ K_{sp} &= 4S^3 \end{aligned} \tag{7.14}$$

To obtain the thermodynamic properties for the dissolution of borax, values for the molar solubility and the solubility product for borax are determined over a range of temperatures. The natural log of K_{sp} is plotted versus $\frac{1}{T}$ to determine the standard enthalpy change, $\Delta H°$, the standard entropy change, $\Delta S°$, and ultimately to calculate the standard free energy change, $\Delta G°$, for the dissolution of borax. See equations 7.9 and 7.7, respectively.

DATA PREP: LINEARIZING THE DATA

In science, straight lines are easier to analyze than curves. The relationship between K_{sp} and Temperature is exponential (a curve).

- **The Trick:** By plotting the **natural log** ($\ln K$) versus the **inverse temperature** ($1/T$), we mathematically "flatten" the curve into a straight line.
- **The Trap:** Because we plot $1/T$, the x-axis is "reversed." High temperatures (large T) appear on the **left** side of the graph (small $1/T$). Keep this in mind when interpreting your slope.

Key Equations

Gibbs Free Energy:

$$\Delta G° = \Delta H° - T\Delta S° \tag{7.15}$$

Gibbs Free Energy and the Equilibrium Constant:

$$\Delta G° = -RT \ln K \tag{7.16}$$

Van't Hoff Equation:

$$\ln K_{sp} = -\frac{\Delta H°}{R}\left(\frac{1}{T}\right) + \frac{\Delta S°}{R} \tag{7.17}$$

7.1: Thermodynamics of Borax - Experiment
7.2: Thermodynamics of Borax - Pre-lab
7.3: Thermodynamics of Borax - Data and Report

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7.1: Thermodynamics of Borax - Experiment

SAFETY PRECAUTIONS

- Hydrochloric acid (HCl) is corrosive; avoid skin contact and inhalation.

EQUIPMENT AND CHEMICALS NEEDED

Table 7.1.1: Materials List

Equipment	Chemicals
Beakers	Borax
50-mL buret	Anhydrous sodium carbonate, Na_2CO_3
5-mL volumetric pipette and pipette pump	0.20 M hydrochloric acid, $HCl(aq)$
Erlenmeyer flask with rubber stopper	methyl orange indicator solution
Temperature baths: • Beaker • Temperature probe • Ice or hot plate	bromocresol green indicator solution

EXPERIMENTAL PROCEDURE

Part A: Standardize the HCl Solution

A standardized HCl solution is prepared using anhydrous sodium carbonate as a primary standard. Sodium carbonate samples of known mass are transferred to Erlenmeyer flasks, dissolved in deionized water, and titrated to a methyl orange endpoint (pH range 3.1 to 4.4) with the prepared hydrochloric acid solution.

$$CO_3^{2-}(aq) + 2\,H_3O^+(aq) \rightarrow 3\,H_2O(l) + CO_2(g) \tag{7.1.1}$$

The flask is heated to near boiling, close to the stoichiometric point of the analysis, to remove the carbon dioxide gas produced in the reaction.

1. **Obtain starting materials.** Get about 2-3 g of anhydrous sodium carbonate, Na_2CO_3, from the desiccator. Get about 200 mL of ~0.20 M HCl.
2. **Prepare the primary standard. Before coming to the lab**, calculate the mass of sodium carbonate that neutralizes 17 mL of 0.20 M HCl at the stoichiometric point [see equation 7.1.1]. Measure this mass to the nearest milligram and transfer it to a 125-mL Erlenmeyer flask. Prepare at least three samples of sodium carbonate for the analysis of the HCl solution.
3. **Prepare the buret.** Clean a buret and rinse it with several 3- to 5-mL portions of the ~0.20 M HCl solution. Use a clean funnel to fill the buret with the ~0.20 M HCl solution. After 10–15 seconds, use the proper technique to read and record the initial volume.
4. **Titrate the primary standards.** To each solid sodium carbonate sample, add ~50 mL water and methyl orange. Dispense HCl while swirling; near the endpoint, heat to boiling to remove CO_2. Continue dropwise to endpoint.
5. **Do the calculations.** Calculate the molarity of the standardized HCl. Trials must agree within 1 %.

Part B: Prepare the Borax Solutions

1. **Calibrate test tubes.** Pipette 5 mL water into six tubes, mark meniscus, discard water, air-dry.
2. **Prepare stock borax solution.** Add 30–35 g borax to 100 mL water, stopper, agitate.
3. **Prepare temperature-dependent test solutions:**
   a. Label six clean test tubes.
   b. Half-fill with saturated borax, place in baths at ~0, room, 32, 40, 48, 56 °C (≤60 °C).

VERNIER TECH BOX: SAMPLING AT EQUILIBRIUM

1. Temperature Monitoring

- Connect the **Temperature Probe** to **CH 1**.
- Place the probe in your borax solution. Ensure the tip is suspended in the liquid, **not touching the bottom** (which may be hotter or colder than the solution).
- Watch the live readout on the screen. You do not need to record a graph.

2. The "Critical Moment" Technique

- **Goal:** You need the concentration of the *dissolved* ions at exactly the measured temperature.
- **Procedure:**
  1. Stir the solution vigorously to ensure saturation.
  2. Stop stirring and let the solid settle for 10–15 seconds.
  3. **The Read:** Record the temperature instantly.
  4. **The Draw:** Immediately decant 5.00 mL of the **clear liquid at the top** (supernatant) into your calibrated tube.

4. **Prepare saturated solutions.**
   a. Agitate for 10–15 min ensuring solid remains present.
   b. Allow to settle and reach thermal equilibrium; record temperatures.
5. **Prepare samples for analysis.** Transfer 5 mL clear supernatant to calibrated test tubes — **no solid borax**.

Part C: Analyze the Borax Test Solutions

1. **Transfer sample.** If crystals form, warm until dissolved; transfer to Erlenmeyer flask with warm rinses.
2. **Titrate the sample.** Dilute to ~25 mL, add bromocresol green, titrate with standardized HCl to yellow endpoint.
3. **Repeat** for all test tubes.

DATA SKILL: DETECTING OUTLIERS

Do not clean your glassware yet!

Before performing any calculations, perform a "Visual Reality Check" on your raw data:

- **The Concept:** Borax is *more* soluble at higher temperatures. Therefore, the volume of HCl required to titrate the solution **must decrease** as the temperature of the sample decreases.
- **The Check:** Sketch a quick plot of **Volume of HCl (y-axis)** vs. **Temperature (x-axis)**.
- **The Decision:** If you see a point that violates this trend (e.g., the volume goes *up* at a lower temperature), that is a chemically impossible outlier. You likely titrated too fast or transferred solid crystals. **Repeat that specific temperature trial immediately.**

CHEMICAL DISPOSAL

- Do not pour chemical waste down the sink. Dispose of waste solutions as instructed by your instructor or lab supervisor.

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7.2: Thermodynamics of Borax - Pre-lab

Name: _____ Date: _____

1. Why is sodium carbonate used to standardize the hydrochloric acid?

2. What mass of sodium carbonate is needed to neutralize 17 mL of 0.20 M HCl?

3. Why is it important to always have solid borax in the test tubes in the various water baths?

4. **Slope Prediction:** The dissolution of Borax is an endothermic process (ΔH° is positive).
   - Look at the Van't Hoff equation: $\ln K = -\frac{\Delta H^\circ}{R}\left(\frac{1}{T}\right) + \frac{\Delta S^\circ}{R}$.
   - Since ΔH° is positive and R is positive, will the slope of your line ($m = -\frac{\Delta H^\circ}{R}$) be **positive** or **negative**?
   - If your Excel plot gives you a slope with the opposite sign, what math error might you have made in your calculations?

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7.3: Thermodynamics of Borax - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Table **7.3.1**: Part A: Standardization of HCl

Trial	
Mass of Na_2CO_3 (g)	
Initial buret reading (mL)	
Final buret reading (mL)	
Volume of HCl needed (mL)	
Molar concentration of HCl (M)	

Table **7.3.2**: Part C: Analysis of Borax Test Solutions

Test Tube	
Approximate Temperature (°C)	
Actual Temperature (°C)	
Initial buret reading (mL)	
Final buret reading (mL)	
Volume of HCl used (mL)	

POST-LAB QUESTIONS

Show a sample calculation for each part. Use a spreadsheet to make and analyze the plot. Include the spreadsheet in your submission.

1. From the titrations, calculate the molarity of the saturated borax solutions at each of the temperatures. These concentrations will be the solubility, *S*, of borax at each of the measured temperatures.

2. Calculate the solubility product, K_{sp}, of borax at each of the measured temperatures ($K_{sp} = 4S^3$).

Table 7.3.3: Results Table

Test tube		
Temperature (°C)		
Molar solubility of borax, S (M)		
$\ln K_{sp}$ of borax		

DATA SKILL: LOGARITHMS IN EXCEL

In Question 3, you will plot the **Natural Logarithm** of K_{sp}. Here are two common pitfalls to avoid:

1. **Negative Values are Normal:** Since K_{sp} is usually a small number (less than 1), the natural log ($\ln K_{sp}$) will be a **negative number**. Do not delete the negative sign!
2. **The "Math Error":** If Excel gives you a `#NUM!` error, check your K_{sp} calculation. You cannot take the log of a negative number or zero. This usually means you subtracted where you should have multiplied.
3. **Function Check:** Ensure you use the Excel function `=LN()` for natural log, not `=LOG()` which is base-10. Using the wrong log base will make your slope (and your ΔH) incorrect by a factor of 2.303.

3. Use Excel to plot the natural logarithm of the solubility product, $\ln K_{sp}$ (y-axis), versus the reciprocal temperature, $\frac{1}{T}$ (x-axis), for each sample and add a Linear Trendline to the graph. Insert an image of the graph here showing the best-fit equation and **Coefficient of Determination (R^2)**.

DATA SKILL: INTERPRETING R^2

When you add a linear trendline in Excel, check the **Coefficient of Determination (R^2)**.

- **What it means:** R^2 ranges from 0 to 1. It tells you how well your data fits the Van't Hoff equation.
- **The Standard:** For this thermodynamic experiment, a careful chemist typically achieves an $R^2 > 0.95$.
- **Analysis:**
  - If $R^2 > 0.98$: Your assumption that ΔH is constant over this temperature range is strongly supported by your data.
  - If $R^2 < 0.90$: The points are "scattered." This suggests random error (titration precision) or that the temperature wasn't stable when you sampled. **You must account for this low reliability in your Conclusion.**

4. From the slope, *m*, of the linear plot, calculate the standard enthalpy of solution of borax, $\Delta H^\circ = -mR$.

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CHAPTER OVERVIEW

8: Redox Reactions

PRE-LAB PREPARATION

*Before taking your **Lab Entry Ticket**, ensure you have mastered the following.*

1. Prerequisite Math & Theory

- **Oxidation States:** You must be able to calculate the oxidation number of Manganese in various compounds.
  - MnO_4^-: Mn is +7 (Purple)
  - MnO_4^{2-}: Mn is +6 (Green)
  - MnO_2: Mn is +4 (Brown solid)
  - Mn^{2+}: Mn is +2 (Colorless/Pink)
- **Redox Definitions:** Remember "LEO says GER" or "OIL RIG."
  - **Oxidation:** Loss of electrons (Charge goes UP).
  - **Reduction:** Gain of electrons (Charge goes DOWN).
  - **Oxidizing Agent:** The substance that *causes* oxidation (it gets reduced).

2. Required Technical Skills

- **Microscale Reaction:** You will use a well plate. Do not overfill the wells. Mixing is done by gently tapping the plate or using a toothpick.
- **Observation:** "Clear" is not a color (it describes transparency). Use "Colorless" for water-like solutions.

3. Critical Safety

- **Strong Oxidizers:** $KMnO_4$ and H_2O_2 are strong oxidizers. They can cause burns and stains.
- **Strong Bases:** 50% $NaOH$ is extremely caustic. Wear gloves and goggles at all times.

PURPOSE

- To investigate the principles of redox reactions by observing the chemical conversion of iron(II) ions to iron(III) ions using various oxidizing agents (such as hydrogen peroxide, potassium permanganate, and sodium hypochlorite) and the reduction of iron(III) ions to iron(II) ions using various reducing agents (such as sodium sulfite, sodium bromide, and sodium iodide).
- To utilize chemical indicators (potassium thiocyanate and potassium ferricyanide) to test for and confirm the presence of iron(III) ions and iron(II) ions, respectively, after redox reactions have occurred.
- To observe and document the characteristic colorful states of manganese (such as Mn^{2+}, MnO_2, and MnO_4^-) corresponding to various oxidation numbers (+2 through +7) and use these visual changes to determine the products of redox reactions and write corresponding balanced half-reactions and overall equations.

INTRODUCTION

1. **Iron Ion Redox Chemistry (Parts A and B)**: These experiments involve observing the conversion between iron(II) ions and iron(III) ions. Part A examines the oxidation of Fe^{2+} to Fe^{3+} using various oxidizing agents, including hydrogen peroxide (H_2O_2), potassium permanganate ($KMnO_4$), and sodium hypochlorite ($NaClO$). Conversely, Part B investigates the reduction of Fe^{3+} to Fe^{2+} using reducing agents like sodium sulfite (Na_2SO_3), sodium bromide ($NaBr$), and sodium iodide (NaI). The presence of the respective iron ions is confirmed using specific chemical indicators: potassium thiocyanate ($KSCN$) is utilized to test for Fe^{3+}, and potassium ferricyanide ($K_3[Fe(CN)_6]$) is used to test for Fe^{2+}.
2. **Colorful States of Manganese (Part C)**: This section focuses on the relationship between oxidation number and visible color for manganese. Reactions are performed involving manganese compounds (such as manganese(II) sulfate and potassium permanganate) to observe and document the characteristic colors corresponding to various oxidation numbers of manganese,

which range from +2 (e.g., Mn^{2+}, colorless or light pink) up to +7 (e.g., MnO_4^-, very dark purple). These observations are then used to analyze and write balanced half-reactions and overall equations for the redox processes.

Manganese occurs in the following oxidation numbers.

Table **8.1**: Oxidation numbers of manganese

Oxidation number	Species	Color
+2	$\mathbf{Mn^{2+}}$ (*aq*)	Colorless or light pink
+3	$\mathbf{Mn^{3+}}$ (*aq*)	Rose or red
+4	$\mathbf{MnO_2}$ (*s*)	Dark brown
+5	$\mathbf{MnO_4^{3-}}$ (*aq*)	Blue
+6	$\mathbf{MnO_4^{2-}}$ (*aq*)	Green
+7	$\mathbf{MnO_4^-}$ (*aq*)	Very dark purple

DATA PREP: PREDICTING COLORS

In redox chemistry, color is your data. Manganese is unique because its color tells you its oxidation state instantly.

- **The Prediction:** Before you mix anything in Part C, look at Table 1. If you start with purple $\mathbf{MnO_4^-}$ (+7) and reduce it slightly to green $\mathbf{MnO_4^{2-}}$ (+6), you know exactly one electron was transferred.
- **The Data:** If you see brown precipitate, you have hit the +4 state ($\mathbf{MnO_2}$). This "visual voltmeter" allows you to determine the number of electrons transferred without doing any math.

Key Concepts

Oxidation vs. Reduction:

- **Oxidation:** Loss of electrons (Oxidation Number increases).
- **Reduction:** Gain of electrons (Oxidation Number decreases).

Oxidizing vs. Reducing Agents:

- **Oxidizing Agent:** The species that *takes* electrons (it is reduced).
- **Reducing Agent:** The species that *gives* electrons (it is oxidized).

8.1: Redox Reactions - Experiment
8.2: Redox Reactions - Pre-lab
8.3: Redox Reactions - Data and Report

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8.1: Redox Reactions - Experiment

SAFETY PRECAUTIONS

- Hydrochloric acid, sulfuric acid, and sodium hydroxide are caustic. Handle with care.
- 50% NaOH is VERY CAUSTIC and must be handled with care. Use gloves.

EQUIPMENT AND CHEMICALS NEEDED

Table **8.1.1**: Materials List

Equipment	Chemicals	Chemicals
6-well plate	0.1 M ammonium iron(II) sulfate, $(NH_4)_2Fe(SO_4)_2$	0.2 M sodium bromide, $NaBr$
10-mL graduated cylinder	3 % hydrogen peroxide, H_2O_2	5 % sodium hypochlorite, $NaClO$
Toothpicks (for stirring)	0.1 M iron(III) chloride, $FeCl_3$	0.2 M sodium iodide, NaI
	0.1 M potassium ferricyanide, $K_3[Fe(CN)_6]$	0.2 M sodium sulfite, Na_2SO_3
	0.1 M potassium thiocyanate, $KSCN$	3 M hydrochloric acid, HCl
	0.006 M potassium permanganate, $KMnO_4$	6 M sulfuric acid, H_2SO_4
	0.025 M potassium permanganate, $KMnO_4$	6 M sodium hydroxide, $NaOH$
	0.12 M manganese(II) sulfate, $MnSO_4$	50 % sodium hydroxide, $NaOH$

EXPERIMENTAL PROCEDURE

For this lab, place a reaction plate on a clean piece of paper and label its wells as shown below. The top wells will be referred to as wells A1, A2, and A3. The bottom wells are B1, B2, and B3.

Part A: React Iron(II) ions with Oxidizing Agents

1. Use a clean 6-well plate and a clean 10-mL graduated cylinder, add 5 mL of 0.1 M ammonium iron(II) sulfate in well A1.
2. Add 5 mL of 0.1 M iron(III) chloride, $FeCl_3$, to well A2.
3. **Record** the initial colors in the A wells in Table A.
4. Add 5 drops of 0.1 M potassium thiocyanate, $KSCN$, to both wells A1 and A2.
5. **Record** the final colors in the A wells in Table A.
6. Add 5 mL of 0.1 M ammonium iron(II) sulfate to wells B1, B2, and B3.
7. Add 20 drops of 3 M HCl to wells B1 and B2.
8. Add 20 drops of 3% hydrogen peroxide to well B1.
9. Add 3 mL of 0.025 M potassium permanganate, $KMnO_4$, to well B2. Stir if needed.
10. Add 3 mL of 5 % sodium hypochlorite, $NaClO$, to well B3. Stir if needed
11. **Record** the initial colors in the B wells in Table A.
12. Test for the presence of iron(III) in wells B1, B2, and B3 by adding 5 drops of 0.1 M potassium thiocyanate.
13. Record the final colors in the B wells in Data Table A.
14. Clean and rinse wells for use in part B.

Part B: React of Iron(III) with Reducing Agents

1. Add 5 mL if 0.1 M ammonium iron(II) sulfate to well A1 and 5 mL of iron(III) chloride into A2.
2. **Record** the initial colors in the A wells in Table B.
3. Add 5 drops of 0.1 M potassium ferricyanide, $K_3[Fe(CN)_6]$ to wells A1 and A2
4. **Record** the final colors in the A wells in Table B.
5. Add 5 mL of 0.1 M iron(III) chloride to wells B1, B2, and B3.

6. Add 20 drops of 3 M **HCl** and 20 drops of 0.2 sodium sulfite, Na_2SO_3, solution to well B1.
7. Add 20 drops of 0.2 M sodium bromide, **NaBr**, to well B2.
8. Add 20 drops of 0.2 M sodium iodide, **NaI**, to well B3.
9. **Record** the initial colors in the B wells in Table B.
10. Test for the presence of iron(II) in wells B1, B2, and B3 by adding 5 drops of 0.1M potassium ferricyanide.
11. **Record** the final colors in the B wells in Table B.
12. Clean and rinse wells for use in part C.

Part C: Observe Colorful States of Manganese

DATA SKILL: PREDICTING COLORS

Use color as data.

- **Purple (MnO_4^-):** Oxidation state +7.
- **Green (MnO_4^{2-}):** Oxidation state +6.
- **Brown Solid (MnO_2):** Oxidation state +4.
- **Colorless (Mn^{2+}):** Oxidation state +2.

By observing the color change, you can determine exactly how many electrons were transferred per manganese atom.

1. Using a clean 10-mL graduated cylinder, add 5 mL of 0.12 M manganese(II) sulfate, $MnSO_4$, solution to 3 wells of the large well plates (label A1 A2 A3).

TECHNIQUE ALERT: COLOR AS DATA

How to count electrons with your eyes:

- In Part C, you are mixing Permanganate (Purple, +7) with Sulfite (Reducing agent).
- **Scenario 1:** If it turns **Green**, the Mn gained 1 electron (+7 → +6).
- **Scenario 2:** If it turns **Brown (muddy)**, the Mn gained 3 electrons (+7 → +4).
- **Scenario 3:** If it turns **Colorless**, the Mn gained 5 electrons (+7 → +2).
- *Your observation of color IS your measurement of the reaction stoichiometry.*

2. **Record** the initial appearance/color of the solution in the A wells in Table C.
3. Add 5 mL of 6 M sulfuric acid to well A2.
4. Add 2–3 drops of the 0.006 M potassium permanganate solution to well A2 using a disposable pipette and stir with a toothpick.
5. Using the same pipette, add 2–3 drops of the 0.006 M potassium permanganate to well A3 and do NOT stir.
6. **Record** the final appearance/color of the solutions in the A wells in Table C.
7. Using a clean 10-mL graduated cylinder, add 5 mL of 0.006 M potassium permanganate solution to the remaining 3 wells of the large well plate (label B1 B2 B3).
8. **Record** the initial appearance/color of the solution in the B wells in Table C.
9. Add 5 mL of 6 M sodium hydroxide to well B2.
10. Using a disposable pipette, add about 2 mL of 0.20 M sodium sulfite to well B2 and stir.
11. Using a clean 10-mL graduated cylinder, add 5 mL of 50 % **NaOH** (VERY CAUSTIC, handle with care) to well B3 and stir rapidly until the solution turns color.
12. **Record** the final colors of the solutions in the B wells in Table C.

CHEMICAL DISPOSAL

- Dispose of all chemicals according to your instructor.

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8.2: Redox Reactions - Pre-lab

Name: _____ Date: _____

1. How can potassium thiocyanate be used to confirm that iron(II) ions have been oxidized to iron(III) ions?

2. Identify at least two different oxidizing agents that are used in Part A.

3. How can potassium ferricyanide be used to confirm that iron(III) ions have been reduced to iron(II) ions?

4. The sulfite ion, used as a reducing agent in Part B, is oxidized to the sulfate ion. Assign oxidation numbers to the sulfur atom in both of these.

5. **Logic Check:** You have three halogens: Chlorine (Cl_2), Bromine (Br_2), and Iodine (I_2).
   - If Cl_2 can oxidize Br^- to Br_2...
   - And Br_2 can oxidize I^- to I_2...
   - **Prediction:** Will Cl_2 be able to oxidize I^-? Why or why not? (This is the "Transitive Property" of redox strength).

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8.3: Redox Reactions - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Table 8.3.1: Part A. Reactions of Iron(II) with Oxidizing Agents

Well	Reactants	Initial color	Final color
A1	Fe^{2+}		
A2	Fe^{3+}		
B1	$Fe^{2+} + H^{+} + H_2O_2$		
B2	$Fe^{2+} + H^{+} + MnO_4^{-}$		
B3	$Fe^{2+} + OCl^{-}$		

Table 8.3.2: Part B. Reactions of Iron(III) with Reducing Agents

Well	Reactants	Initial color	Final color
A1	Fe^{2+}		
A2	Fe^{3+}		
B1	$Fe^{3+} + H^{+} + SO_3^{2-}$		
B2	$Fe^{3+} + Br^{-}$		
B3	$Fe^{3+} + I^{-}$		

Table 8.3.3: Part C. Colorful Oxidation Numbers of Manganese

Well	Reactants	Initial color	Final color
A1	Mn^{2+}		
A2	$Mn^{2+} + H^{+} + MnO_4^-$		
A3	$Mn^{2+} + MnO_4^-$		
B1	MnO_4^-		
B2	$MnO_4^- + SO_3^{2-} + OH^-$		
B3	$MnO_4^- + OH^-$		

POST-LAB QUESTIONS

Questions for Parts A and B

1. Use the oxidation number rules to assign oxidation numbers for the indicated atoms in each oxidizing agent and its product (Part A).

Atom	Oxidizing Agent	Oxidation number	Product	Oxidation number
Mn	MnO_4^-		Mn^{2+}	
O	H_2O_2		H_2O	
Cl	ClO^-		Cl^-	

2. Fill in the blanks to show the number of electrons involved in each half-reaction (balance charge).
   a. $MnO_4^-(aq) + 8H^+(aq) + ___ e^- \longrightarrow Mn^{2+}(aq) + 4H_2O(l)$
   b. $H_2O_2(aq) + 2H^+(aq) + ___ e^- \longrightarrow 2H_2O(l)$
   c. $ClO^-(aq) + H_2O(l) + ___ e^- \longrightarrow Cl^-(aq) + 2OH^-(aq)$

3. Combine the oxidation half-reaction for Fe^{2+} to Fe^{3+} with the appropriate half-reaction from Question #2 and write the balanced equation for the overall redox reaction of Fe^{2+} with:
   a. Permanganate ion, MnO_4^-

   b. Hydrogen peroxide, H_2O_2

   c. Hypochlorite ion, ClO^-

4. a. Fill in the blank to show the number of electrons in the following half-reaction:

b. Write the balanced equation for the overall redox reaction of Fe^{3+} with the sulfite ion.

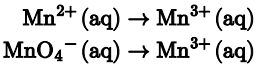
5. Based on your observations in Part B, which halide — bromide ion or iodide ion — is the stronger reducing agent? Explain.

Questions for Part C

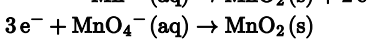
For each of the following wells:

1. Explain the color of the product.
2. Balance the half-reactions.
3. Write the overall balanced equation for the reaction.

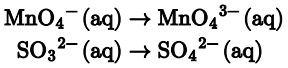
Well A2: Potassium permanganate is added to an acidified solution of manganese(II) sulfate.



Well A3: Drops of potassium permanganate are added to an acidified solution of manganese(II) sulfate.



Well B2: A solution of sodium sulfite is added to a basic potassium permanganate solution.



Well B3: A highly basic sodium hydroxide solution is added to a solution of potassium permanganate.

$$\mathrm{MnO_4^-\,(aq) \rightarrow MnO_4^{3-}\,(aq)}$$
$$\mathrm{OH^-\,(aq) \rightarrow O_2\,(g)}$$

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CHAPTER OVERVIEW

9: Voltaic Cells

PRE-LAB PREPARATION

*Before taking your **Lab Entry Ticket**, ensure you have mastered the following.*

1. Prerequisite Math & Theory

- **Cell Potential:** $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$.
  - Remember: The substance with the **higher (more positive)** reduction potential wants to be the Cathode.
- **Concentration Cells (Nernst):** Even if both electrodes are Copper, a voltage exists if concentrations differ.

$$E = E^\circ - \frac{0.0592}{n}\log Q \tag{9.1}$$

Since $E^\circ = 0$ (same metal), the voltage is driven entirely by $Q = \frac{[Dilute]}{[Concentrated]}$.

2. Required Technical Skills (See Appendix 13.1)

- **Section A (Voltage Probe):** Connecting the probe. (Note: The Voltage Probe uses the same channel as pH/Temp).
- **Technique:** Polishing electrodes with steel wool to remove oxide layers is critical for good contact.

3. Critical Safety

- **Heavy Metals:** Solutions containing Copper, Zinc, and Unknown metals must be disposed of in the Heavy Metal Waste container.

PURPOSE

- To prepare a Cu/Zn voltaic cell and measure the potential.
- To test two voltaic cells that use unknown metal electrodes and identify the metals.
- To prepare a copper concentration cell and observe and measure its potential.

INTRODUCTION

In electrochemistry, a voltaic cell is a specially prepared system in which an oxidation-reduction reaction occurs spontaneously. This spontaneous reaction produces an easily measured electrical potential. Voltaic cells have a variety of uses.

In this experiment, you will prepare a variety of semi-microscale voltaic cells in a 24-well test plate. A voltaic cell is constructed by using two metal electrodes and solutions of their respective salts (the electrolyte component of the cell) with known molar concentrations. In Parts I and II of this experiment, you will use a Voltage Probe to measure the potential of a voltaic cell with copper and zinc electrodes. You will then test two voltaic cells that have unknown metal electrodes and, through careful measurements of the cell potentials, identify the unknown metals. In Part III of the experiment, you will measure the potential of a special type of voltaic cell called a concentration cell. In the concentration cell, you will observe how a voltaic cell can maintain a spontaneous redox reaction with identical copper metal electrodes, but different electrolyte concentrations.

DATA PREP: CONCENTRATION AS A BATTERY

Usually, a battery is driven by two *different* metals. In Part C, you use Copper for both sides. The driving force is **Entropy**.

- Nature wants to equalize the concentrations. Electrons flow from the dilute side to the concentrated side to balance them out.
- **The Nernst Prediction:** Since $E^\circ = 0$ (same metal), the entire voltage comes from the term $-\frac{0.0592}{n}\log Q$. If your measured voltage is 0.00 V, you likely put the same solution in both wells!

Key Equations

Standard Potential:

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} \tag{9.2}$$

Nernst Equation at 25 °C:

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0592\ \text{V}}{n} \log Q \tag{9.3}$$

9.1: Voltaic Cells - Experiment

9.2: Voltaic Cells - Pre-lab

9.3: Voltaic Cells - Data and Report

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9.1: Voltaic Cells - Experiment

SAFETY PRECAUTIONS

- Wear safety goggles at all times.

EQUIPMENT AND CHEMICALS NEEDED

Table **9.1.1**: Materials List

Equipment	Chemicals
voltage probe and leads	0.10 M $Cu(NO_3)_2$
24-well plate	0.10 M $Zn(NO_3)_2$
string	copper strips
steel wool	zinc strips
disposable pipettes	1 M KNO_3
	0.10 M unknown X; unknown X metal strips
	0.10 M unknown Y; unknown Y metal strips

EXPERIMENTAL PROCEDURE

Part A: Determine the *E*° for a Cu/Zn Voltaic Cell

1. Use a 24-well test plate. Transfer ~1 mL each of 0.10 M $Cu(NO_3)_2$ and 0.10 M $Zn(NO_3)_2$ into two adjacent wells.
2. Obtain copper and zinc metal strips. Polish with steel wool. Connect the Voltage Probe leads and prepare to place electrodes in their respective solutions.

VERNIER TECH BOX: MEASURING VOLTAGE

1. Setup

- Connect the **Voltage Probe** to **CH 1**.
- The screen will display the live potential in Volts (V).

2. Connecting the Clips (The Polarity Rule)

- **RED Clip:** Connect to the metal you think is the **Cathode** (Reduction/High Potential).
- **BLACK Clip:** Connect to the metal you think is the **Anode** (Oxidation/Low Potential).

3. Interpreting the Sign

- **Positive Voltage (+):** You guessed correctly! Electrons are flowing into the Red clip.
- **Negative Voltage (-):** You guessed backwards. The Black clip is actually the Cathode. Swap the clips to get a positive reading.

3. Make a salt bridge by soaking a short string in 1 M KNO_3.
4. Connect the wells with the salt bridge, ensuring it contacts both solutions.
5. Connect voltmeter leads: positive (red) to the cathode, negative (black) to the anode.
6. Measure the potential of the Cu/Zn cell.
   a. Insert electrodes and record the stabilized potential. If negative, switch the leads and re-measure.
   b. Remove, clean, and polish both electrodes.
   c. Re-insert and record a second potential.
   d. Clean both electrodes again.

e. Re-insert and record a third potential.
f. Calculate and record the mean potential.

Part B: Determine the *E*° for Two Voltaic Cells Using Copper and Unknown Metals

1. Obtain 0.10 M unknown X and its metal strip.
2. Transfer a small amount of 0.10 M X to a well adjacent to $Cu(NO_3)_2$.
3. Prepare a new KNO_3 salt bridge and connect the wells.
4. Measure the X/Cu cell potential three times and average as in Part A.
5. Repeat for unknown Y to determine the Y/Cu cell potential.

DATA SKILL: THE LOGIC MATRIX

To determine the activity series of your unknowns, look at the polarity of your voltmeter:

- **If Red (Cathode) is on Copper:** Copper is being reduced. The Unknown Metal is the Anode (being oxidized). Therefore, the Unknown is a **stronger reducing agent** than Copper.
- **If Red (Cathode) is on Unknown:** The Unknown is being reduced. Copper is the Anode. Therefore, the Unknown is a **weaker reducing agent** than Copper.

Use this logic to rank X, Y, Zn, and Cu from strongest to weakest.

Part C: Prepare and Test a Copper Concentration Cell

1. Prepare 20 mL of 0.050 M $CuSO_4$ by mixing 1 mL of 1.0 M $CuSO_4$ with 19 mL water.
2. Set up a concentration cell with 0.050 M $CuSO_4$ in one well and 1.0 M $CuSO_4$ in the adjacent well. Use copper electrodes and a KNO_3 salt bridge.
3. Measure and record the potential as done previously.

CHEMICAL DISPOSAL

- Dispose of all chemicals according to your instructor.

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9.2: Voltaic Cells - Pre-lab

Name: _____ Date: _____

1. Write the balanced equation and calculate the theoretical standard cell potential, E°_{cell}, for the following cell:

$$Zn\,(s)\ |\ Zn^{2+}\,(aq)\ ||\ Cu^{2+}\,(aq)\ |\ Cu\,(s)$$

2. Use the table of standard reduction potentials in your text, or another approved reference, to complete the following table. An example is provided.

Table **9.2.1**: Galvanic Cells

Electrodes (circle the anode)	**Reduction Half-reactions**	**Standard Reduction Potential, E° [V]**	**Voltaic Cell Potential, E° [V]**
(Zn) Cu	$Zn^{2+} + 2\,e^- \longrightarrow Zn\,(s)$ $Cu^{2+} + 2\,e^- \longrightarrow Cu\,(s)$	−0.76 +0.34	+1.10
Cu **Ag**			
Cu **Mg**			
Zn **Sn**			
Cu **Sn**			

3. For the Zn/Cu cell, if the concentration of Zn^{2+} is 0.01 M and the concentration of Cu^{2+} is 1.0 M, will the measured E_{cell} be greater than, less than, or equal to the E°_{cell}? Explain your reasoning based on the Nernst equation.

4. **Logic Check:** You measure a cell potential of $+0.46$ V for an Unknown Metal X connected to Copper ($+0.34$ V).
   - If Copper is the cathode (Red), what is the reduction potential of X?

- If Copper is the anode (Ox), what is the reduction potential of X?

- Look at a standard potential table. Which scenario matches a real metal (Ag, Mg, Zn, Sn)?

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9.3: Voltaic Cells - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Table 9.3.1: Data Table

	Part A	Part B		Part C
Description	Cu/Zn	X/Cu	Y/Cu	Cu Concentration Cell
Which electrode is the anode?				
First Measurement (V)				
Second Measurement (V)				
Third Measurement (V)				
Average cell potential (V)				

POST-LAB QUESTIONS

1. **(Part A)** Compare the average cell potential for your Cu/Zn cell with the E°_{cell} that you calculated in the pre-lab exercise. Calculate your **percent error** and **explain** why your cell potential is different from the text value.

2. **(Part B)** The unknown metals X and Y were either magnesium, silver, tin, or zinc. Use the text value for the reduction potential of copper and the measured cell potentials for the unknowns to identify X and Y.

3. **(Part C)** Use the Nernst equation to calculate the theoretical value of the potential, E, of the copper concentration cell and compare this value with the cell potential that you measured. Assume a temperature of 298 K.

DATA SKILL: VALIDATING THE MODEL

You calculated a theoretical E using the Nernst equation. Now compare it to your real voltmeter reading.

- Calculate the **Percent Difference**.
- **Analysis:** If your measured voltage is significantly *lower* than theoretical, what does that imply about the true concentration of your "0.050 M" solution? (Hint: Did you dilute it perfectly?)

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CHAPTER OVERVIEW

10: Decomposition of Hydrogen Peroxide

PRE-LAB PREPARATION

*Before taking your **Lab Entry Ticket**, ensure you have mastered the following.*

1. Prerequisite Math & Theory

- **Method of Initial Rates:** You will determine the reaction order (n) by comparing two runs where the concentration changes.

$$\frac{rate_2}{rate_1} = \left(\frac{[KI]_2}{[KI]_1}\right)^n \tag{10.1}$$

  - *Example:* If doubling the concentration doubles the rate (2^1), it is First Order.
- **Arrhenius Equation:** To find Activation Energy (E_a), you will compare rates at two temperatures (Part A vs. Part D).

$$\ln\left(\frac{k_2}{k_1}\right) = \frac{E_a}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right) \tag{10.2}$$

2. Required Technical Skills (See Appendix 13.1)

- **Section G:** Setting Time Duration and Rate (you need 300 seconds).
- **Section F (Linear Fit):** Finding the slope of the "Initial Rate" region.
- **Data Skill:** Selecting only the straight/linear portion of the graph (the first 60 seconds) to avoid error from reagent depletion.

3. Critical Safety

- **Pressure Hazard:** The reaction produces Oxygen gas. If the stopper is stuck, pressure can build up. **Never** point the test tube at yourself or others.

PURPOSE

- To conduct the catalyzed decomposition of hydrogen peroxide under various conditions.
- To determine the rate law expression for the reaction.
- To calculate the rate constant for the reaction.
- To calculate the activation energy for the reaction.

INTRODUCTION

The decomposition of hydrogen peroxide, $\mathbf{H_2O_2}$, in an aqueous solution proceeds very slowly. For instance, a commercial bottle of 3 % hydrogen peroxide is stable for a long duration. The decomposition reaction is described by the equation:

$$2\,H_2O_2\,(aq) \rightarrow 2\,H_2O\,(l) + O_2\,(g) \tag{10.3}$$

To accelerate this naturally slow process, several catalysts can be utilized, including potassium iodide (KI), manganese(IV) oxide ($\mathbf{MnO_2}$), or the enzyme catalase.
In this experiment, the catalyzed decomposition of hydrogen peroxide is carried out in a closed vessel. By doing so, the reaction rate can be determined by monitoring the pressure increase that occurs as oxygen gas ($\mathbf{O_2}$) is produced.

This approach allows for several calculations and determinations:

- The rate constant for the reaction can be calculated.
- By varying the initial molar concentration of the hydrogen peroxide solution (as detailed in Parts A, B, and C), the rate law expression for the reaction can be determined.

- By conducting the reaction at different temperatures (such as approximately 20 °C and 30 °C, as detailed in Part D), the activation energy (E_a) for the reaction can be calculated.

DATA PREP: THE INITIAL RATE METHOD

As the reaction proceeds, the concentration of $\mathbf{H_2O_2}$ decreases, causing the reaction to slow down. This means your Pressure vs. Time graph will eventually curve.

- **The Goal:** We want the rate at the *very beginning*, where concentrations are known (initial mixing).
- **The Method:** You must fit a linear trendline only to the **first 60 seconds** (or the straightest early region). If you include the curved part later in the run, your slope (rate) will be artificially low.

Key Equations

Rate Law:

$$rate = k[\mathrm{H_2O_2}]^m[\mathrm{I^-}]^n \tag{10.4}$$

Initial Rate:

$$rate = \frac{\Delta[\text{reactant}]}{\Delta t} \tag{10.5}$$

Determining Order:

$$\frac{rate_1}{rate_2} = \left(\frac{[\text{reactant}]_2}{[\text{reactant}]_1}\right)^n \tag{10.6}$$

10.1: Decomposition of Hydrogen Peroxide - Experiment
10.2: Decomposition of Hydrogen Peroxide - Pre-lab
10.3: Decomposition of Hydrogen Peroxide - Data and Report

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10.1: Decomposition of Hydrogen Peroxide - Experiment

SAFETY PRECAUTIONS

- Pressure will build up in the test tube. Carefully remove the stopper from the test tube to relieve the pressure.

EQUIPMENT AND CHEMICALS NEEDED

Table **10.1.1**: Materials List

Equipment	Chemicals
Gas pressure sensor with data-logging capability	H_2O_2 (3 % hydrogen peroxide)
Tubing	0.5 M KI
Thermometer or temperature probe	
One-hole rubber stopper with stem	
Solid rubber stopper	
Medium test tube	
Two 10 mL graduated cylinders	
Graduated plastic pipette	
1 L beaker	

EXPERIMENTAL PROCEDURE

Part A: Decompose 3 % H_2O_2 solution with 0.5 M KI solution at ~20 °C

1. Prepare the reagents for temperature equilibration.
   a. Obtain room-temperature water to set up a water bath to completely immerse the test tube. Use a thermometer or a temperature probe to measure the temperature of the bath. Record this temperature in your data table for Parts A–C; presume that the water bath temperature remains constant throughout.
   b. Add 4 mL of 3 % H_2O_2 solution to the test tube. Seal the test tube with the solid rubber stopper and place the test tube in the water bath.
   c. Measure 2 mL of 0.5 M KI solution in a graduated cylinder. Draw 1 mL of the KI solution into a graduated pipette. Invert the pipette and immerse the reservoir end of the pipette in the water bath.
2. Use the tubing to connect the one-hole rubber stopper to the gas pressure sensor.

VERNIER TECH BOX: KINETICS SETUP

1. Sensor Setup

- Connect the **Gas Pressure Sensor** to **CH 1**.
- Ensure the rubber stopper is clean and dry. A wet stopper may slide out under pressure.

2. Time Profile Configuration (Critical)

- Tap the **Rate** box (usually says "Time Based").
- **Duration:** Change to **300 s** (5 minutes).
- **Interval:** Change to **0.5 s/sample** (or 2 samples/s).
- Tap **OK**.

3. Zeroing the Sensor?

- **Do NOT zero the sensor.** We are measuring the *change* in pressure (slope), so the starting value (atmospheric pressure) doesn't matter.

3. Set the gas pressure data-collection rate to 0.1 samples/second and the data-collection length to 300 seconds.
4. Start data collection. Data will be collected for five minutes.
5. Run the reaction and collect pressure data. **Quickly do the following:**
   a. Remove the test tube from the water bath and remove the solid stopper.
   b. Remove the plastic pipette from the water bath and quickly transfer the 1 mL of KI solution into the test tube. Tap or lightly shake the test tube to mix the reagents.
   c. Seal the test tube with the one-hole stopper connected to the gas pressure sensor.
   d. Place the test tube back in the water bath.
   e. If necessary, gently hold the test tube so that it stays completely immersed in the water bath.
6. When the data collection is complete, carefully remove the stopper from the test tube to relieve the pressure.
7. Determine the initial rate of the reaction.
   a. Identify a **linear** region just beyond the initial flat portion of the graph.
   b. Select about one minute of data starting from the beginning of the linear region.
   c. Fit this selected data to a linear curve.

> DATA SKILL: CHECK THE FIT
>
> **Look at your linear fit.**
>
> - Your R^2 value should be > **0.98**.
> - If R^2 is low, you likely included the "flat" part at the start (before mixing) or the "curved" part at the end (depletion).
> - **Fix:** Adjust your selection brackets to highlight *only* the straightest, steepest part of the slope.

   d. Record the slope, in your data table, as the initial rate for the reaction.
8. Clean the test tube for the next part.

Part B: Decompose 3 % H_2O_2 solution with 0.25 M KI solution at ~20 °C

1. Add 4 mL of 3 % H_2O_2 solution into the test tube. Seal the test tube with the solid rubber stopper and place the test tube in the water bath.
2. Add 1 mL of deionized water to the remaining 1 mL of KI solution in the graduated cylinder. Swirl the mixture gently to mix the solution.
3. Draw 1 mL of the KI solution into a plastic pipette. Invert the pipette and immerse the reservoir end of the pipette in the water bath. Allow the test tube and the pipette to remain in the water bath for at least two minutes before proceeding.
4. Repeat Steps 4–8 from Part A.

Part C: Decompose 1.5 % H_2O_2 solution with 0.5 M KI solution at ~20 °C

1. Prepare a 1.5 % H_2O_2 solution by mixing 2 mL of deionized water with 2 mL of 3 % H_2O_2 solution. Transfer all 4 mL of the 1.5 % H_2O_2 solution to the test tube, seal the test tube with the solid stopper, and place the test tube in the water bath.
2. Rinse and clean the graduated cylinder that you have used for the KI solution. Add a fresh 2 mL of 0.5 M KI solution to the graduated cylinder.
3. Draw 1 mL of the KI solution into a plastic pipette. Invert the pipette and immerse the reservoir end of the pipette in the water bath. Allow both the test tube and the pipette to remain in the water bath for at least two minutes before proceeding.
4. Repeat Steps 4–8 from Part A.

Part D: Decompose 3 % H_2O_2 solution with 0.5 M KI solution at ~30 °C

1. Conduct Part D identically to the procedure in Part A, with one exception: use a water bath at around 30 °C.

CHEMICAL DISPOSAL

- Dispose of the contents of the test tube as directed.

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10.2: Decomposition of Hydrogen Peroxide - Pre-lab

Name: _____ Date: _____

1. Briefly explain the role of a catalyst in a chemical reaction. How does using a catalyst affect the reaction's **activation energy** and its **initial rate**? [Image of reaction coordinate diagram showing catalyst effect]

2. The decomposition reaction is: $2\,H_2O_2\,(aq) \longrightarrow 2\,H_2O\,(l) + O_2\,(g)$. Why does the pressure build up in the closed vessel?

3. The hydrogen peroxide solution that you are using in this experiment is labeled as a 3 % solution, mass/volume (% g/mL). This means that there are 0.03 g H_2O_2 per mL of solution. However, to complete the calculations, the concentration must be in molarity. Calculate the molarity of a 3 % mass/volume H_2O_2 solution (Part A, B, and C) and a 1.5 % mass/volume H_2O_2 solution (Part D) and record these values in the table below. Report your answers using 2 significant figures.

Table **10.2.1**: Solution Volumes and Molarities

Part	Volume H_2O_2 (mL)	$[H_2O_2]$ *before* mixing	Volume KI (mL)	[KI] *before* mixing
A	4		1	0.50 M
B	4		1	0.25 M
C	4		1	0.50 M
D	4		1	0.50 M

Show your calculations here:

4. **Unit Conversion Check:** Your pressure sensor measures the rate in **kPa/s**, but the rate law requires **M/s** (mol/L·s).
   Rearrange the Ideal Gas Law ($PV = nRT$) to solve for $\frac{n}{V}$ (Molarity).
   Show that:

$$\text{Rate } [\text{M/s}] = \frac{\text{Rate } [\text{kPa/s}]}{R \times T} \tag{10.2.1}$$

   (Hint: Watch your units for R! If pressure is in kPa, which R value must you use?)

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10.3: Decomposition of Hydrogen Peroxide - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Table **10.3.1**: Data Table

Part	Reactants	Temperature (°C)	Initial rate (kPa/s)
A	4 mL 3.0 % H_2O_2 + 1 mL 0.50 M KI		
B	4 mL 3.0 % H_2O_2 + 1 mL 0.25 M KI		
C	4 mL 1.5 % H_2O_2 + 1 mL 0.50 M KI		
D	4 mL 3.0 % H_2O_2 + 1 mL 0.50 M KI		

Attach your four pressure vs. time curves here, showing the initial rate (kPa/s) determination (the linear fit).

POST-LAB QUESTIONS

Fill in the following Results Table. Instructions are below.

Table **10.3.1**: Results Table

Part	Initial rate (kPa/s)	Initial rate (M/s)	$[H_2O_2]$ *after* mixing	$[I^-]$ *after* mixing	Rate constant, k (Fill in units)
A					
B					
C					
D					

1. You need to convert pressure in kPa to molarity of O_2 so that the rate is in units of M/s. **Use the ideal gas law**. Show the calculation for Part A here:

2. Use the dilution equation to calculate the initial concentrations of H_2O_2 and I^- after mixing. Show the calculation for Part A here:

3. Write the rate law expression for the catalyzed decomposition of hydrogen peroxide. Instead of guessing the reaction order by looking at the rates, calculate the exact value of the exponent n using logarithms.

   Starting with the ratio of rates: $\frac{\text{Rate}_2}{\text{Rate}_1} = \left(\frac{[\text{Reactant}]_2}{[\text{Reactant}]_1}\right)^n$

   Take the log of both sides and rearrange to solve for n:

$$n = \frac{\log\left(\frac{\text{Rate}_2}{\text{Rate}_1}\right)}{\log\left(\frac{[\text{Reactant}]_2}{[\text{Reactant}]_1}\right)} \tag{10.3.1}$$

   Your calculated n will likely be a decimal (e.g., 0.92 or 2.15). Round this calculated value to the nearest whole number (0, 1, or 2) to determine the true kinetic order.

DATA SKILL: ROUNDING REACTION ORDERS

Real data is messy.

- When you calculate the exponent n using logarithms, you might get a number like 0.92 or 1.15.
- **The Rule:** Reaction orders usually are integers (0, 1, or 2) for simple kinetic steps.
- **Action:** Round your calculated decimal to the nearest whole number to write your final Rate Law.
  - $0.8 - 1.2 \rightarrow$ Order = 1
  - $1.8 - 2.2 \rightarrow$ Order = 2

4. Calculate the rate constant, *k*, for each part, using the rate law that you determined. Be sure to use the correct units for *k*. Show your work for Part A here:

5. Use the two-point Arrhenius equation (shown below) to determine the activation energy, E_a, for this reaction. Use $R = 8.314\ \text{J mol}^{-1}\text{K}^{-1}$ and express your answer in kJ mol^{-1}.

$$\ln\left(\frac{k_1}{k_2}\right) = -\frac{E_a}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right) \tag{10.3.2}$$

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CHAPTER OVERVIEW

11: Determination of an Activation Energy

PRE-LAB PREPARATION

*Before taking your **Lab Entry Ticket**, ensure you have mastered the following.*

1. Prerequisite Math & Theory

- **The Modified Arrhenius Equation:** Because we measure *time* (t) instead of rate (k), the slope equation flips sign:

$$\ln t = \frac{E_a}{R}\left(\frac{1}{T}\right) + \text{constant} \tag{11.1}$$

  - **Slope:** $m = +\frac{E_a}{R}$. (Positive slope means longer time at lower temperature).
  - **Calculation:** $E_a = \text{Slope} \times 8.314$.
- **Catalyst Effect:** A catalyst lowers the activation energy. Therefore, your calculated E_a for Part B (Catalyzed) must be **lower** than Part A (Uncatalyzed).

2. Required Technical Skills (See Appendix 13.1)

- **Section A:** Measuring Temperature (without logging).
- **Data Skill:** Linearizing data ($\ln t$ vs $1/T$) in Excel.

3. Critical Safety

- **Oxidizer Hazard:** Ammonium Peroxydisulfate is a strong oxidizer. Avoid contact with skin.
- **Waste:** All solutions contain Iodine and Copper. Dispose in the **Heavy Metal Waste** container.

PURPOSE

- **To investigate the effect of temperature and a catalyst on the rate of a chemical reaction** by studying the oxidation of iodide (I^-) to iodine (I_2) by peroxydisulfate ($S_2O_8^{2-}$). The time required for a given amount of reaction will be measured at various temperatures, both with and without a catalyst.
- **To calculate the activation energy (E_a)** for the reaction by measuring the reaction rate (inversely proportional to time, t) at different absolute temperatures (T) and applying the logarithmic form of the Arrhenius equation: $\ln t = \frac{E_a}{R}\left(\frac{1}{T}\right) + \text{constant}$.
- **To determine the effect of a catalyst** on the reaction's activation energy (E_a) and compare the E_a value obtained for the uncatalyzed reaction with the value obtained when a Cu^{2+} catalyst is used.
- **To graphically analyze the experimental data** by plotting the natural logarithm of time ($\ln t$) versus the reciprocal of the absolute temperature ($1/T$) to find the slope, which is then used to calculate the activation energy, (E_a).

INTRODUCTION

As noted in your previous experiment, two of the factors that affect the rate of a chemical reaction are **temperature** and a **catalyst**. In this experiment, we will investigate in more detail just how those factors affect the rate of a chemical reaction.

From considerations of kinetic theory, we know that only a small proportion of the many molecular collisions result in a reaction. In most collisions, the molecules simply rebound like billiard balls, without change. Only the more energetic molecules collide with sufficient energy for their electron orbitals to interpenetrate, thus forming an **activated complex**. These then react to form more stable products. The proportion of such energetic molecules depends upon the absolute temperature and is in accord with the Maxwell-Boltzmann distribution law, which defines the increasing proportion of such more energetic molecules at higher temperatures. The minimum net energy that the colliding molecules must possess to react is called the **activation energy, E_a**. The situation is somewhat analogous to an attempt to roll a ball up an incline, over a hump, and down to a lower, more stable level. The extra energy necessary to get over the hump is the activation energy. The activation energy for a chemical reaction can be found experimentally by measuring the reaction rate at different temperatures.

A **catalyst** is a substance that permits the formation of the activated complex more easily, with lower activation energy. At a given temperature, a larger proportion of the molecules will possess enough energy to react, and thus the catalyst increases the rate of the reaction.

How can we calculate the activation energy? The relationship between the reaction rate constant, *k*, the absolute temperature, *T*, and the activation energy, E_a, is given by the equation

$$k = A \exp\left(-\frac{E_a}{RT}\right) \tag{11.2}$$

Where

R is the gas constant (8.314 J/mol·K),

A is a constant for the reaction.

The factor $\exp\left(-\frac{E_a}{RT}\right)$ comes from the Boltzmann principle, which states that, for the equilibrium at any absolute temperature *T*, the ratio of the number of molecules with large energy to the number with small energy is a function of the factor $\exp\left(-\frac{\Delta E}{RT}\right)$, where ΔE is the difference in energies. We may rewrite equation 11.2 in the logarithmic form

$$\ln k = -\frac{E_a}{R}\left(\frac{1}{T}\right) + \ln A \tag{11.3}$$

Since the rate constant, *k*, is inversely proportional to the time, *t*, for a given amount of reaction, we may rewrite equation 11.3 as

$$\ln \frac{c}{t} = -\frac{E_a}{R}\left(\frac{1}{T}\right) + \ln A \tag{11.4}$$

where *c* is some constant. Manipulating this equation gives

$$\ln c - \ln t = -\frac{E_a}{R}\left(\frac{1}{T}\right) + \ln A \tag{11.5}$$

or

$$\ln t = \frac{E_a}{R}\left(\frac{1}{T}\right) + constant \tag{11.6}$$

Thus, if we **graph the natural log of time ($\ln t$) against the reciprocal of the absolute temperature ($1/T$)**, the slope, *m*, of the straight line thus obtained, multiplied by *R* will give us the activation energy E_a.

$$m = \frac{E_a}{R} \quad \text{or} \quad E_a = mR \tag{11.7}$$

In this experiment, you will determine the activation energy, both with and without a catalyst, for the ***oxidation*** of iodide (I^-) to iodine (I_2) by peroxydisulfate ($S_2O_8^{2-}$), also known as persulfate. In the process, the peroxydisulfate is ***reduced*** to sulfate (SO_4^{2-}).

$$I^-(aq) + S_2O_8^{2-}(aq) \rightarrow I_2(aq) + 2SO_4^{2-}(aq) \tag{11.8}$$

This reaction is rather slow, and you can easily measure the time required for a given amount of reactants by having a large excess of iodide and peroxydisulfate and a small amount of thiosulfate ($S_2O_3^{2-}$), with starch as an indicator. Thiosulfate reacts rapidly to ***reduce*** the iodine back to iodide. In the process, the thiosulfate is ***oxidized*** to tetrathionate ($S_4O_6^{2-}$).

$$I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq) \tag{11.9}$$

Reaction 11.9 ensures that no iodine remains in the solution until the thiosulfate is exhausted. At this point, the iodine formed from reaction 11.8 suddenly turns the starch blue.

DATA PREP: TIME VS. RATE CONSTANT

The Arrhenius equation normally uses the rate constant k. However, in this lab, you measure time (t).

- **The Math:** Since rate is inversely proportional to time ($k \propto 1/t$), we can substitute $\ln(1/t) = -\ln(t)$.

- **The Slope:** This means the slope of your $\ln(t)$ vs $1/T$ graph will be **positive** (E_a/R), whereas a standard Arrhenius plot ($\ln k$ vs $1/T$) has a **negative** slope ($-E_a/R$).
- **Sanity Check:** If your slope is negative, check your axis labels!

Key Equations

Linear Arrhenius Equation:

$$\ln k = -\frac{E_a}{R}\left(\frac{1}{T}\right) + \ln A \tag{11.10}$$

Activation Energy (for this experiment):

$$slope = \frac{E_a}{R} \quad \text{or} \quad E_a = slope \times R \tag{11.11}$$

11.1: Determination of an Activation Energy - Experiment
11.2: Determination of an Activation Energy - Pre-lab
11.3: Determination of an Activation Energy - Data and Report

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11.1: Determination of an Activation Energy - Experiment

SAFETY PRECAUTIONS

- **Chemical Hazards:** Peroxydisulfate salts ($S_2O_8^{2-}$) are strong oxidizing agents and should be handled with care. Copper(II) salts are toxic. Standard laboratory personal protective equipment (PPE) including gloves, goggles, and lab coats must be worn.
- **Temperature Hazards:** Caution should be exercised when handling the warm water bath and ice bath to prevent burns or frostbite.
- **Ingestion/Contact:** Avoid contact with skin and eyes, and do not ingest any of the chemicals.

EQUIPMENT AND CHEMICALS NEEDED

Table 11.1.1: Materials List

Equipment	Chemicals
Burets	0.200 M potassium iodide, KI
50-mL Erlenmeyer flasks (2)	0.100 M ammonium peroxydisulfate, $(NH_4)_2S_2O_8$
250-mL beaker	0.0050 M sodium thiosulfate, $Na_2S_2O_3$
400-mL Beaker	Starch solution
Temperature Probes or Thermometers	0.1 M copper(II) sulfate, $CuSO_4$
Timer	
Ice Bath	
Hot Water Bath	

EXPERIMENTAL PROCEDURE

Part A: Determine the Effect of Temperature

Carry out rate measurements at different temperatures as follows. In all cases, keep the concentration the same. For all mixtures, you will use:

- 20.00 mL of 0.200 M potassium iodide, KI
- 20.00 mL of 0.100 M ammonium peroxydisulfate, $(NH_4)_2S_2O_8$
- 10.00 mL of 0.0050 M sodium thiosulfate, $Na_2S_2O_3$
- 3 drops of starch solution

Table 11.1.2: Part A Experimental Conditions

Mixture 1:	1. **Fill three burets**, one with 0.200 M **KI**, one with 0.100 M $(NH_4)_2S_2O_8$, and 0.0050 M $Na_2S_2O_3$ into the third. 2. Carefully measure exactly 20.00 mL of iodide and peroxydisulfate solutions **separately into two labeled 50-mL Erlenmeyer flasks**, which have been rinsed with distilled water and drained. 3. **Rinse and drain a 250-mL beaker** to be used as the reaction vessel. Add exactly 10.00 mL of 0.0050 M $Na_2S_2O_3$ and 3 drops of starch indicator to the reaction beaker, and place a temperature probe in the beaker. **VERNIER TECH BOX: TEMPERATURE MONITORING** **1. Setup** ○ Connect the **Temperature Probe** to **CH 1**. ○ You do **not** need to set up a graph or hit "Collect." Just watch the live meter readout. **2. The Mixing Technique** ○ **Initial T:** Read the temperature of the solutions *before* mixing. ○ **Final T:** Read the temperature the instant the color changes. ○ **Calculation:** Use the **Average Temperature** for your Arrhenius plot. 4. Have your timer ready, and **quickly pour the iodide and peroxydisulfate solutions into the beaker and swirl** gently to promote mixing. Read the initial temperature. 5. Time the reaction from mixing to the moment the starch ***begins*** to change color, and read the final temperature. 6. Record the average of the two temperatures. 7. Repeat the rate measurement with a second set of samples, which should check the first data **within about 1 second**.
Mixture 2:	1. **Measure** out the solutions as Mixture 1. 2. **Warm** the flasks in warm water so that the final reaction temperature is about **30–35 °C**. 3. The 250-mL reaction beaker may also be warmed in a 400-mL beaker of slightly warm water. 4. **Mix** the solutions, **time** the reaction, and read the **temperature**.
Mixture 3:	1. Follow Mixture 2, but adjust to obtain a final reaction temperature of **45–50 °C**.
Mixture 4:	1. Measure out the solutions as Mixture 1. 2. Cool the flasks to **10–15 °C** using an ice-water bath. 3. Cool the reaction beaker similarly if needed. 4. Mix, time, and record the temperature.
Mixture 5:	1. Follow Mixture 4, but cool further to reach a final reaction temperature of **3–7 °C**.

DATA SKILL: REALITY CHECK

Before you clean up, check your raw data trend.

- **The Concept:** Reaction rates increase with temperature. This means the reaction should finish *faster* as it gets hotter.
- **The Check:** Look at your timer values. As Temperature increases, Time must **decrease**.
- **The Trap:** If your 45°C trial took *longer* than your 35°C trial, you have an error (solution preparation or temperature measurement). Repeat that trial immediately.

Part B: Determine the Effect of a Catalyst

1. Empty the peroxydisulfate solution from its buret into a 250-mL beaker and add enough stock peroxydisulfate to make 150 mL total.
2. Add 3 drops of 0.1 M $\mathbf{CuSO_4}$ (providing $\mathbf{Cu^{2+}}$) as a catalyst and mix thoroughly.
3. Refill the buret with this catalyzed peroxydisulfate solution.

Table **11.1.3**: Part B Experimental Conditions

Mixture 6:	1. Carry out a rate measurement using: ◦ 20.00 mL of the peroxydisulfate–$\mathbf{Cu^{2+}}$ solution ◦ 20.00 mL of iodide solution ◦ 10.00 mL of thiosulfate solution ◦ 3 drops of starch at room temperature
Mixtures 7–10:	1. Measure out solutions as for Mixture 6, but perform reactions at **two higher** and **two lower** temperatures (similar to Mixtures 2–5).

CHEMICAL DISPOSAL

- Dispose of all chemicals according to your instructor.

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11.2: Determination of an Activation Energy - Pre-lab

Name: _____ Date: _____

1. To determine the activation energy (E_a), what two values are plotted against each other, and how is the slope of the resulting straight line used in conjunction with the gas constant (R) to calculate E_a? Write down the relevant equation.

2. If the slope of an Arrhenius plot is -6.4×10^3 K, calculate the activation energy, E_a, in kJ/mol.

3. In the reaction between iodide and peroxydisulfate, describe the role of thiosulfate and starch in making the reaction time-dependent and visually measurable?

4. What is the visual signal that marks the end of the timed reaction?

5. **Graphical Prediction:** You will plot two lines on the same graph: Uncatalyzed and Catalyzed.
   - Which line should have a steeper slope? (Recall: Steeper slope = Higher E_a).
   - Which line should be lower on the y-axis (representing faster times/smaller t) at the same temperature?

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11.3: Determination of an Activation Energy - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Table 11.3.1: Data Table

Mixture Number	Average Temperature, T (°C)	Time, t (s)
1		
2		
3		
4		
5		
6		
7		
8		
9		
10		

POST-LAB QUESTIONS

Use a spreadsheet to do the following

1. Convert temperature to Kelvin.
2. Calculate the $\ln t$ and $\frac{1}{T}$ for each mixture.
3. Use a spreadsheet to plot these values with $\ln t$ on the y-axis and $\frac{1}{T}$ (in K^{-1}) on the x-axis.
4. Perform a linear least-squares fit on the data.
5. Find the slope of the fit and use it to calculate the activation energy, E_a, from equation 11.6.
6. Calculate the $\ln t$ and $\frac{1}{T}$ (in K^{-1}) for catalyzed mixtures 6–10 and plot them **on the same graph** as the uncatalyzed mixtures.
7. Determine E_a for the catalyzed mixtures as before.

DATA SKILL: THE UNIT TRAP

Watch your units carefully!

- Your slope has units of **Kelvin (K)**.
- The gas constant R is $8.314\ \text{J/mol} \cdot \text{K}$.
- When you multiply $\text{Slope} \times R$, your result is in **Joules per mole (J/mol)**.
- **The Fix:** You must divide by 1000 to report your answer in **kJ/mol** as requested in the table.

Table **11.3.2**: Results Table

	uncatalyzed	catalyzed
slope (units= _____)		
intercept (units = _____)		
E_a (kJ/mol)		

Insert your graphs here:

8. From your measurements on the **uncatalyzed** mixtures 1–5, comment on the statement, “The rate of a chemical reaction doubles for every 10 °C temperature increase.”

   Don't just say "yes" or "no." Test the rule with your actual data.

   - **The Test:** Calculate the ratio of time at **20°C** vs **30°C** (or any 10-degree gap).
   - **Ratio** $= \frac{t_{low}}{t_{high}}$
   - **Analysis:** Is this ratio close to 2.0? If it is 1.5 or 3.0, the "Rule of Thumb" is not accurate for this specific reaction.

9. Describe how the catalyst affects the rate of a chemical reaction.

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CHAPTER OVERVIEW

12: Radioactivity Simulation

PRE-LAB PREPARATION

*Before taking your **Lab Entry Ticket**, ensure you have mastered the following.*

1. Prerequisite Math & Theory

- **First-Order Decay:** Radioactive decay follows the same math as First-Order Kinetics.

$$\ln N_t = -\lambda t + \ln N_0 \tag{12.1}$$

  - **y-axis:** $\ln$(Dice Remaining)
  - **x-axis:** Round Number (Time)
  - **Slope:** $-\lambda$ (Decay Constant)
- **Half-Life Math:** $t_{1/2} = \frac{\ln 2}{\lambda}$.
  - If you know the slope (λ), you can calculate how many rounds it takes for half the dice to disappear.

2. Required Skills

- **Simulation Logic:** Understand that rolling 10 dice 10 times is statistically equivalent to rolling 100 dice once.
- **Data Skill:** Graphing $\ln N$ vs. Round Number in Excel to find the decay constant.

3. Critical Safety

- **Choking Hazard:** Dice are small. Do not put them in your mouth.
- **Slip Hazard:** Keep dice contained in a box; loose dice on the floor can cause falls.

PURPOSE

- To simulate radioactive decay by studying a random process, specifically rolling dice.
- To better understand the concept of half-life ($t_{1/2}$) by observing how the number of dice remaining decreases with each round.
- To determine the "half-life" of the dice by graphing the number of dice remaining after each round.
- To study how a spontaneous and completely random process (radioactive decay) can be simulated using the probability of a specific outcome in another random process (like rolling a die on a "decay number").

INTRODUCTION

Radioactive decay is a spontaneous and completely random process. There is no way to predict how long it will take a specific atom of a radioactive isotope to disintegrate and produce a new atom. The probability, however, that a particular atom will decay after a certain period can be simulated by studying other random processes, such as a coin toss or a “roll of the dice.”

Radioactive nuclei disintegrate through different processes at varying rates. The time required for different radioactive nuclei to decay varies widely, from seconds or minutes for highly unstable nuclei to billions of years or more for long-lived radioactive nuclei. Polonium-218, for example, emits alpha particles and decays very quickly − within minutes. Uranium-238 also decays via alpha-particle production, but the decay takes place over billions of years! The relative rate of decay of different radioactive isotopes is most conveniently described by comparing their half-lives. The half-life ($t_{1/2}$) of a radioactive isotope (called a radioisotope) is the amount of time needed for one-half of the atoms in a sample to decay. Every radioisotope has a characteristic half-life, which is independent of the total number of atoms in a sample. Thus, the half-life of polonium-218 is about three minutes, while the half-life of uranium-238 is more than 4 billion years. Regardless of the total number of atoms in a sample of polonium-218, one-half of the atoms will always “disappear” (decompose to produce other atoms) within three minutes.

To better understand half-life, consider the following example. Iodine-131 is an artificial radioisotope of iodine that is produced in nuclear reactors for use in medical research and in nuclear medicine. It has a half-life of eight days. If 32 grams of iodine-131 are originally produced in a nuclear reactor, after eight days, only 16 grams of iodine-131 will remain. After two half-lives, or 16 days, only 8 grams will be left, and after three half-lives (24 days), only 4 grams will be left. Every eight days, the amount of iodine-131 that remains will decrease by 50%.

The process of radioactive decay may be modeled by studying a random process such as a coin toss. One could make the rule that a coin landing on 'heads' does not decay, but if the coin lands on 'tails', it decays. Imagine placing 100 coins heads-up in a box to start, shaking the box, and then discarding the coins that land tails or “decay.” Since the probability of a specific coin “decaying” (landing tails) is 50%, we would predict that only 50 coins will remain in the box after the first coin toss. Repeating the coin toss with 50 coins remaining in the box should result in an additional 25 coins “decaying” in the second round, and so on (see the table below). A simulated “radioactive decay curve” obtained by graphing the data (Figure 1) shows that the “half-life” of coins is equal to “one coin toss.” The number of coins remaining in the box decreases by 50% after each coin toss.

Table **12.1**: “Radioactive Decay” of Coins

Round (Coin Toss)	**1**	**2**	**3**	**4**	**5**	**6**	**7**
Number of coins (initial)	100	50	25	12	6	3	1
Number of coins that “decay”	50	25	13	6	3	2	1
Number of coins remaining	50	25	12	6	3	1	0

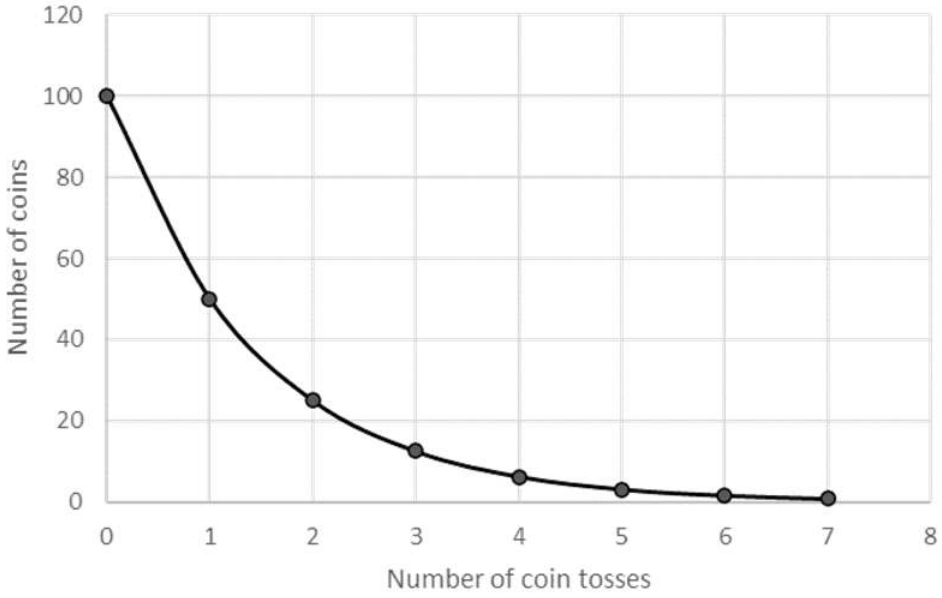


Figure **12.1**: "Radioactive Decay" of coins.

The purpose of this activity is to simulate radioactive decay by studying the probability of a random process – rolling dice. The “radioactive decay” of dice will be studied by rolling 10 dice ten times in Round 1 and recording the number of dice that display a specific “decay number,” for example, all dice that read six. (Rolling 10 dice ten times is equivalent to rolling 100 dice once.) The total number of dice that “decayed” (landed on six) during Round 1 will then be counted and subtracted from the total number of

dice rolled. This is the number of dice remaining that will be rolled in Round 2. This process will be repeated until no dice remain. The “half-life” of dice will be determined by graphing the number of dice remaining after each round.

DATA PREP: KINETICS REVISITED

Mathematically, Radioactive Decay is identical to **First-Order Kinetics** (which you studied in the Hydrogen Peroxide lab).

- **The Rate Law:** Rate = $k[N]$. The rate of decay depends only on the number of atoms (dice) present.
- **The Graph:** Just like in kinetics, plotting the **natural log** of the number of dice ($\ln N$) versus the Round Number (t) yields a straight line. The slope of this line is $-\lambda$ (the decay constant).

Key Equations

First-order, radioactive decay:

$$N_t = N_0 e^{-\lambda t} \quad \text{or} \quad t = -\frac{1}{\lambda}\ln\left(\frac{N_t}{N_0}\right) \tag{12.2}$$

Radioactive decay half-life:

$$\lambda = \frac{\ln 2}{t_{1/2}} \quad \text{or} \quad t_{1/2} = \frac{\ln 2}{\lambda} \tag{12.3}$$

12.1: Radioactivity Simulation - Experiment

12.2: Radioactivity Simulation - Pre-lab

12.3: Radioactivity Simulation - Data and Report

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12.1: Radioactivity Simulation - Experiment

SAFETY PRECAUTIONS

- Dice are a choking hazard and should not be put in your mouth.
- Use a box to contain the dice so that they do not fall on the floor and cause a tripping hazard.

EQUIPMENT AND CHEMICALS NEEDED

Table **12.1.1**: Materials List

Equipment
10 dice (different groups will have different-sided dice)

EXPERIMENTAL PROCEDURE

Part A: Simulate Radioactive Decay with Dice

1. Obtain an assigned set of 10 multi-sided dice (e.g., 4-sided, 6-sided, 8-sided, 10-sided, 12-sided, or 20-sided dice). Your instructor will assign a decay number to each group. Record the number of sides on the dice and the assigned decay number in the Data Table.
2. Roll all 10 dice in a box or on a tabletop. Hint: Roll the dice so they all have an opportunity to have a random roll and all the dice land flat. Re-roll any dice that do not land flat.

TECHNIQUE ALERT: RANDOMNESS IS KEY

How to ensure valid data:

- **The Shake:** Use a cup or box. Shake vigorously for 3 seconds before every roll.
- **The Surface:** Roll onto a flat surface. If a die lands "cocked" (leaning against a wall or another die), **re-roll ONLY that die**.
- **The Count:** Be methodical. If you miss counting a "decayed" die, your population (N_t) stays artificially high, ruining your half-life calculation.

3. Assume that any dice landing on the assigned decay number have decayed. Record the number of dice that decayed in the Data column of the Data Table.
4. Repeat steps 2 and 3 until the 10 dice have been rolled 10 times (for a combined total of 100 dice rolls). After each roll, record the number of dice with the assigned decay number in the Data column of the Data Table.
5. Add together the number of dice that decayed from the Data column. Enter this number − the total number of dice that decayed in the first round − in the “Number of dice that decayed” column of the Data Table.
6. Subtract the number of dice that decayed from the initial number of dice (100 for the first round) to determine the number of dice remaining. Record the “Number of dice remaining” in the Data Table and enter this number as the initial number of dice for the second round.
7. Repeat steps 2–6. Roll the dice as many times as needed to match the number of dice (initial) for Round 2. Record all data in the Data Table. For example, if 79 dice remain after the first round, then the 10 dice would be rolled seven times, and then nine dice would be rolled once so that the total number of dice rolled in Round 2 is equal to 79 [(10×7)+9 = 79].

DATA SKILL: SAMPLE SIZE NOISE

The Law of Large Numbers: Radioactive decay is probabilistic. Predictions work best with huge numbers of atoms.

- **Early Rounds (Many Dice):** Your data will likely match the theoretical curve closely because 100 is a decent sample size.
- **Late Rounds (Few Dice):** As the number of dice drops below 10, your data will get "noisy" (jumpy). One random roll has a huge percentage impact. This simulates why calculating half-lives from tiny samples is difficult.

8. Repeat step 7 until no dice remain or until 20 rounds of decay have been completed.

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12.2: Radioactivity Simulation - Pre-lab

Name: _____ Date: _____

1. Radioactive decay follows first-order kinetics. Write the mathematical relationship between half-life, $t_{1/2}$, and the decay constant, λ.
2. Strontium-90 is a radioactive isotope with a half-life of 29 years. Assume that 10,000 atoms of Sr-90 are generated in a nuclear reaction and then stored.
   a. Complete the following table to determine how much time will have elapsed when 1250 atoms of Sr-90 remain in the storage facility.

Table 12.2.1: Sr-90 Radioactive Decay

Number of half-lives	**0**	**1**	**2**	**3**	**4**	**5**
Number of years	0	29				
Number of Sr-90 atoms	10,000					

   b. Graph the data for the radioactive decay of Sr-90 and draw a continuous, smooth-fit curve through the data points. This can be done in Excel by make an Exponential Trendline. Insert your graph here.

   c. Describe the shape of the radioactive decay curve shown above.

   d. What is the probability that **a specific Sr-90 atom** will have decayed after (a) 29 years and (b) 58 years?

3. **Theoretical Prediction:** Say you use 6-sided dice.
   - If 1 side is "decay" and 5 sides are "safe," what is the probability, P, of decay per roll?
   - The theoretical decay constant is $\lambda_{theo} = -\ln(1-P) = -\ln\left(\frac{\text{Sides}-1}{\text{Sides}}\right)$.
   - Calculate the theoretical half-life: $t_{1/2} = \frac{\ln 2}{\lambda}$.

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12.3: Radioactivity Simulation - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Table **12.3.1**: Data Table

	Number of sides on the dice:			
	Assigned "decay number":			
Round	**Initial Dice**	**Total Decayed**	**Dice Remaining**	**Data (Record decays for each of the 10 rolls)**
0	-	-	100	
1	100			
2				
3				
4				
5				
6				
7				
8				
9				
10				
11				
12				
13				
14				
15				
16				
17				
18				
19				
20				

POST-LAB QUESTIONS

1. Use an Excel spreadsheet to graph the results obtained for the radioactive decay of dice as a scatterplot with symbols only. Note: Include a point on the graph for "100" as the number of dice remaining after zero rolls of the dice (Round zero). Insert

your graph here after completing the steps below:

DATA SKILL: THE LAW OF LARGE NUMBERS

Look at your graph. Is it a perfect smooth curve, or is it "bumpy"?

- **The Cause:** You only used 100 dice. If you flipped a coin 4 times, you might get 3 Heads (75%) instead of 50%. This is **Statistical Noise**.
- **The Fix:** If you rolled 10,000 dice, the points would fall almost perfectly on the trendline. The "bumpiness" is not human error—it is a fundamental property of small sample sizes.

2. Determine the decay constant, λ: Add a Trendline to the graph and select "Exponential", "Display Equation on chart", and "Display R-squared value on chart".

DATA SKILL: EXPERIMENT VS. THEORY

Your Excel trendline gave you an **experimental** λ (the exponent). Now, compare this to the theoretical value for your specific dice.

- **Calculate Theoretical λ:** Use the formula derived from probability:

$$\lambda_{theo} = -\ln\left(\frac{\text{Sides}-1}{\text{Sides}}\right) \tag{12.3.1}$$

- **The Comparison:** Calculate the **Percent Error** between your experimental slope and this theoretical value.
- **Analysis:** If your error is large (>10%), was it because of the small sample size (100 dice), or did you perhaps count the "decay number" incorrectly?

3. Use the fitted decay constant to determine the half-life, $t_{1/2}$.

4. Does the half-life depend on the “decay number” used? Explain, based on probability.

5. Predict how many dice should have remained after 15 rounds using this decay constant. Compare this with the number of dice that actually remained (not your fitted Trendline - your actual data). What factors might account for any difference between the predicted and actual number of dice remaining after 15 rounds?

6. Does the decay curve (your data) for dice have the same general shape as the decay curve for Sr-90 (see the Pre-Lab Questions)? Would the shape of the decay curve be different if you had started with 1000 dice instead of 100?

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CHAPTER OVERVIEW

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13.2.9: Formation Constants for Complex Ions

13.2.10: Standard Electrode (Half-Cell) Potentials

13.2.11: Half-Lives for Several Radioactive Isotopes

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13.1: Vernier LabQuest 3 Instructions

This appendix contains step-by-step instructions for the common tasks used in General Chemistry II. Refer to these sections whenever a lab procedure directs you to "Calibrate" or "Setup" a sensor.

A. Temperature Probe

Measuring Temperature

The Stainless Steel Temperature Probe is auto-identified by the LabQuest.

1. Plug the probe into **CH 1**.
2. The screen will display the live temperature in °C.
3. **Note:** You typically do *not* need to press "Start/Collect" to read a temperature. Just use the live meter readout.
4. **Calibration:** These probes are factory-calibrated and rarely need adjustment. If the reading seems wrong (e.g., room temperature is 80°C), check that the probe is not set to °F or K in the units menu.

B. pH Sensor

Two-Point pH Calibration

pH probes drift over time. For accurate K_a or titration data, you must calibrate before use.

1. Connect the pH Sensor to **CH 1**.
2. Tap the red **pH** box on the meter screen → Select **Calibrate**.
3. Tap **Calibrate Now**.
4. **First Point (Acidic):**
   - Rinse the probe with distilled water and place it in the **pH 4.00** buffer solution.
   - Wait for the voltage reading to stabilize.
   - Enter **4.00** in the "Value 1" field.
   - Tap **Keep**.
5. **Second Point (Neutral):**
   - Rinse the probe and place it in the **pH 7.00** buffer solution.
   - Wait for the voltage to stabilize.
   - Enter **7.00** in the "Value 2" field.
   - Tap **Keep**.
6. Tap **OK**. The probe is now calibrated.

C. Drop Counter (Automated Titration)

Calibrating Drops/mL

The "Drop Counter" counts drops, but you need to know Volume (mL). Because every plastic reagent reservoir tip is slightly different, you must tell the device exactly how large your drops are.

1. Connect the Drop Counter to **DIG 1**.
2. Tap the **Mode** box → Select **Drop Counting**.
3. Tap the red **Drop Count** box → Select **Calibrate** → **Calibrate Now**.
4. **The Procedure:**
   - Place a clean **10-mL graduated cylinder** directly under the drop tip.
   - Open the bottom valve fully to establish a steady, slow drop rate (about 1-2 drops per second).

- Allow the drops to fall until the cylinder is nearly full (between 9 and 10 mL).
- **Close the valve completely.**

5. **The Calculation:**
   - Read the exact volume in the graduated cylinder (e.g., 9.45 mL).
   - Enter this value into the LabQuest.
   - Tap **Keep**.
6. The device now calculates a precise "Drops/mL" ratio for your specific setup.

D. Manual Data Entry

"Events with Entry" Mode

Use this mode when you want to type in one variable (like Concentration or Volume) while the sensor records the other (like Absorbance or pH).

1. Tap the **Mode** box (default is usually "Time Based").
2. Select **Events with Entry** from the dropdown menu.
3. **Configure the X-Axis:**
   - **Name:** Enter the variable you will type (e.g., "Concentration").
   - **Units:** Enter the units (e.g., "mol/L" or "M").
4. Tap **OK**.
5. **How to Collect Data:**
   - Tap **Start/Collect**.
   - When the sensor reading stabilizes, tap **Keep**.
   - A keyboard will appear. Type in the value for your X-axis (e.g., "0.05") and tap **OK**.
   - Repeat for all samples. Tap **Stop** when finished.

E. Spectrometer (SpectroVis Plus)

Calibration & Setup

Step 1: Calibration (Blanking)

1. Allow the lamp to warm up for 2-3 minutes.
2. Fill a cuvette 3/4 full with your blank solvent (usually water).
3. Insert the cuvette so the clear sides face the arrow/light path.
4. Tap the red **Absorbance** box → **Calibrate**.
5. Tap **Finish Calibration** → **OK**. (The reading should be 0.000).

Step 2: Finding λ_{max} (Peak Wavelength)

1. Insert a cuvette containing your most concentrated colored solution.
2. Tap **Start/Collect**. A spectrum graph will appear.
3. Tap **Stop**.
4. Tap the highest point on the curve. The wavelength (nm) and absorbance will appear on the right. Note this wavelength.

Step 3: Setting Up for Beer's Law

1. Tap the **Meter** tab (gauge icon).
2. Tap the **Mode** box.
3. Select **Events with Entry** (See Appendix A.5).
4. The device will now lock onto the λ_{max} you found in Step 2.

F. Linear Fit (Regression)

Finding Slope and Intercept

To find the best-fit line for your data (e.g., Beer's Law or Kinetics):

1. Tap the **Graph** tab.
2. Tap **Analyze** → **Curve Fit**.
3. Select the checkbox for your data series (usually "Absorbance").
4. From the "Fit Equation" dropdown, select **Linear**.
5. The screen will display the equation form $y = mx + b$:
   - **m (Slope):** The rate of change or molar absorptivity.
   - **b (Intercept):** The y-value at zero.
   - **r (Correlation):** A value close to 1.000 indicates a good fit.
6. Tap **OK** to display the line on the graph.

G. Time-Based Settings (Kinetics)

Changing Duration & Rate

For Kinetics or Temperature experiments, you often need to adjust how long the sensor records.

1. Tap the **Meter** tab.
2. Tap the **Mode** box (it usually says "Time Based").
3. **Duration:** Change this to match your reaction time (e.g., change "180 s" to "600 s").
4. **Rate/Interval:**
   - **Rate:** Samples per second (e.g., 2 samples/s).
   - **Interval:** Time between samples (e.g., 0.5 s/sample).
5. Tap **OK**.

H. Managing Data Runs

Storing & Hiding Runs

If you want to do multiple trials (e.g., Trial 1 and Trial 2) without losing your first graph:

1. **Store a Run:** Tap the **File Cabinet Icon** (next to the "Run 1" label). The screen will clear and prepare for "Run 2." Your old data is safe.
2. **Viewing Multiple Runs:**
   - Tap the **Graph** tab.
   - Tap the **Run Number** button (e.g., "Run 2").
   - Select **All Runs** to see them overlayed.
3. **Strike Through Data (Outliers):**
   - Tap the **Table** tab.
   - Tap and drag to select the "bad" data cells.
   - Tap **Table** (menu) → **Strike Through Data**.
   - The graph will ignore these points, but the data remains visible with a line through it.

I. Graph Analysis Tools

Statistics & Tangents

1. Statistics (Mean/Max/Min)

1. On the **Graph** tab, drag across the region of interest.
2. Tap **Analyze** → **Statistics**.
3. Check the box for your data column.
4. A box will appear showing the Mean (Average), Min, and Max for that selection.

2. Tangent (Finding Slope at a Point)

Use this for Titrations to find the steepest part of the curve (Equivalence Point).

1. Tap **Analyze** → **Tangent**.
2. Tap anywhere on the graph curve.
3. A crosshair will appear. The number displayed is the **Slope** at that exact instant.
4. For titrations, the Equivalence Point is where the slope is **highest** (maximum value).

J. Exporting Data

Emailing Files for Excel

To analyze your data in Excel or Google Sheets, you must email the text file to yourself.

1. Ensure the LabQuest is connected to WiFi (Connections icon).
2. Tap **File** → **Email**.
3. **Option A: Text File (Recommended for Excel)**
   - Tap the **Text** icon.
   - Enter your email address and subject.
   - Tap **Send**. (You can open the .txt file in Excel).
4. **Option B: Screenshot (For Lab Reports)**
   - Tap the **Screenshot** icon (Camera).
   - Enter your email and send.

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SECTION OVERVIEW

13.2: Tables

13.2.1: The Periodic Table

13.2.2: Fundamental Physical Constants

13.2.3: Water Properties

13.2.4: Composition of Commercial Acids and Bases

13.2.5: Standard Thermodynamic Properties for Selected Substances

13.2.6: Ionization Constants of Weak Acids

13.2.7: Ionization Constants of Weak Bases

13.2.8: Solubility Products

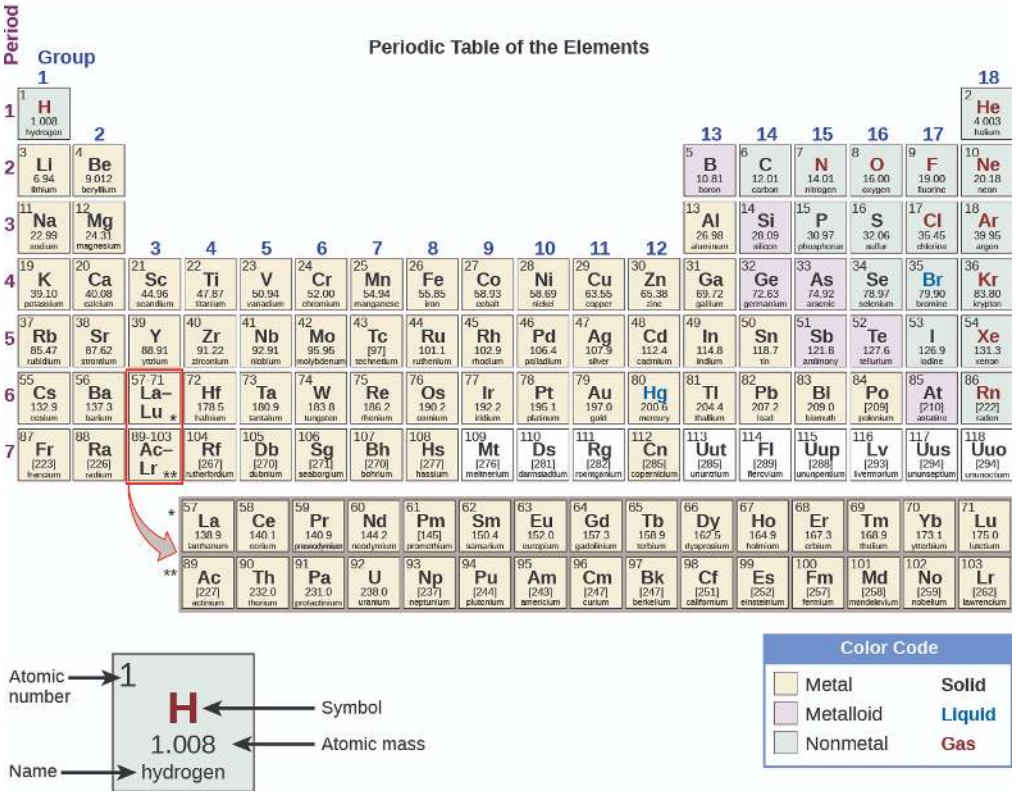
13.2.9: Formation Constants for Complex Ions

13.2.10: Standard Electrode (Half-Cell) Potentials

13.2.11: Half-Lives for Several Radioactive Isotopes

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13.2.1: The Periodic Table



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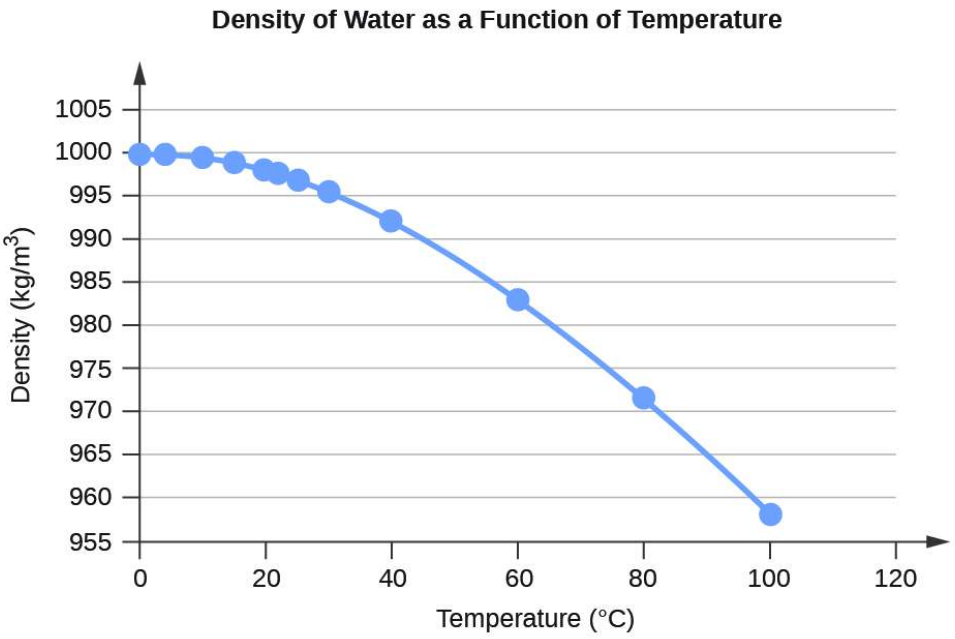
13.2.2: Fundamental Physical Constants

Fundamental Physical Constants	
Name and Symbol	Value
atomic mass unit (amu)	$1.6605402 \times 10^{-27}$ kg
Avogadro's number	6.0221367×10^{23} mol^{-1}
Boltzmann's constant (k)	1.380658×10^{-23} J K^{-1}
charge-to-mass ratio for electron (e/m_e)	$1.75881962 \times 10^{11}$ C kg^{-1}
electron charge (e)	$1.60217733 \times 10^{-19}$ C
electron rest mass (m_e)	$9.1093897 \times 10^{-31}$ kg
Faraday's constant (F)	9.6485309×10^{4} C mol^{-1}
gas constant (R)	8.205784×10^{-2} L atm mol^{-1} K^{-1} = 8.314510 J mol^{-1} K^{-1}
molar volume of an ideal gas, 1 atm, 0 °C	22.41409 L mol^{-1}
molar volume of an ideal gas, 1 bar, 0 °C	22.71108 L mol^{-1}
neutron rest mass (m_n)	$1.6749274 \times 10^{-27}$ kg
Planck's constant (h)	$6.6260755 \times 10^{-34}$ J s
proton rest mass (m_p)	$1.6726231 \times 10^{-27}$ kg
Rydberg constant (R)	$1.0973731534 \times 10^{7}$ m^{-1} = $2.1798736 \times 10^{-18}$ J
speed of light (in vacuum) (c)	2.99792458×10^{8} m s^{-1}

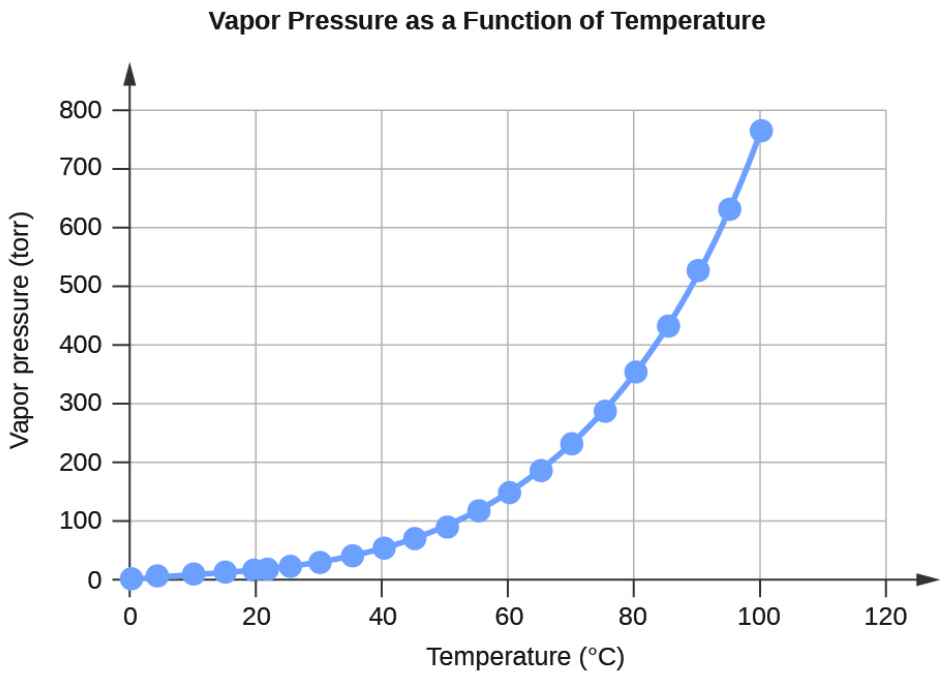
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13.2.3: Water Properties

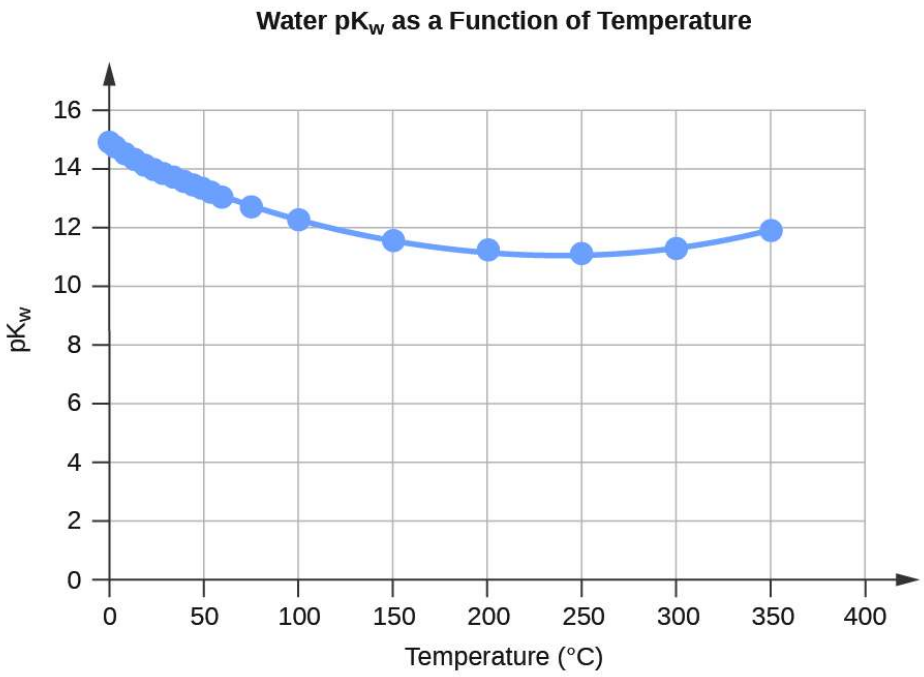
Water Density (kg/m^3) at Different Temperatures (°C)	
Temperature[1]	Density
0	999.8395
4	999.9720 (density maximum)
10	999.7026
15	999.1026
20	998.2071
22	997.7735
25	997.0479
30	995.6502
40	992.2
60	983.2
80	971.8
100	958.4



Water Vapor Pressure at Different Temperatures (°C)		
Temperature	Vapor Pressure (torr)	Vapor Pressure (Pa)
0	4.6	613.2812
4	6.1	813.2642
10	9.2	1226.562
15	12.8	1706.522
20	17.5	2333.135
22	19.8	2639.776
25	23.8	3173.064
30	31.8	4239.64
35	42.2	5626.188
40	55.3	7372.707
45	71.9	9585.852
50	92.5	12332.29
55	118.0	15732
60	149.4	19918.31
65	187.5	24997.88
70	233.7	31157.35
75	289.1	38543.39
80	355.1	47342.64
85	433.6	57808.42
90	525.8	70100.71
95	633.9	84512.82
100	760.0	101324.7



Water K_w and pK_w at Different Temperatures (°C)		
Temperature	K_w 10^{-14}	pK_w[2]
0	0.112	14.95
5	0.182	14.74
10	0.288	14.54
15	0.465	14.33
20	0.671	14.17
25	0.991	14.00
30	1.432	13.84
35	2.042	13.69
40	2.851	13.55
45	3.917	13.41
50	5.297	13.28
55	7.080	13.15
60	9.311	13.03
75	19.95	12.70
100	56.23	12.25

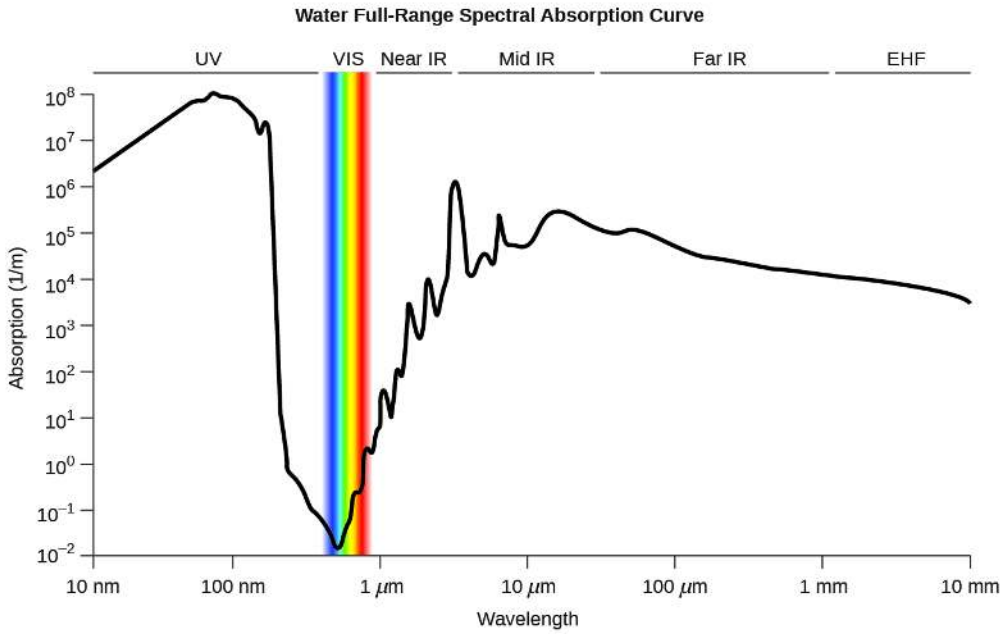


Specific Heat Capacity for Water
C°($H_2O(l)$) = 4184 $J \cdot K^{-1} \cdot kg^{-1}$ = 4.184 $J \cdot g^{-1} \cdot °C^{-1}$
C°($H_2O(s)$) = 1864 $J \cdot K^{-1} \cdot kg^{-1}$
C°($H_2O(g)$) = 2093 $J \cdot K^{-1} \cdot kg^{-1}$

Standard Water Melting and Boiling Temperatures and Enthalpies of the Transitions		
	Temperature (K)	ΔH (kJ/mol)
melting	273.15	6.088
boiling	373.15	40.656 (44.016 at 298 K)

Water Cryoscopic (Freezing Point Depression) and Ebullioscopic (Boiling Point Elevation) Constants
K_f = 1.86°C·kg·mol^{-1} (cryoscopic constant)
K_b = 0.51°C·kg·mol^{-1} (ebullioscopic constant)

<figure class="ui-has-child-figcaption" id="CNX_Chem_00_EE_LiqWatAbso">



<figcaption>Water full-range spectral absorption curve. This curve shows the full-range spectral absorption for water. The *y*-axis signifies the absorption in 1/cm. If we divide 1 by this value, we will obtain the length of the path (in cm) after which the intensity of a light beam passing through water decays by a factor of the base of the natural logarithm e (e = 2.718281828).</figcaption> </figure>

Footnotes

1. 1 Data for t < 0 °C are for supercooled water
2. 2 $pK_w = -\log_{10}(K_w)$

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13.2.4: Composition of Commercial Acids and Bases

Composition of Commercial Acids and Bases			
Acid or Base[1]	Density (g/mL)[2]	Percentage by Mass	Molarity
acetic acid, glacial	1.05	99.5%	17.4
aqueous ammonia[3]	0.90	28%	14.8
hydrochloric acid	1.18	36%	11.6
nitric acid	1.42	71%	16.0
perchloric acid	1.67	70%	11.65
phosphoric acid	1.70	85%	14.7
sodium hydroxide	1.53	50%	19.1
sulfuric acid	1.84	96%	18.0

Footnotes

1. 1 Acids and bases are commercially available as aqueous solutions. This table lists properties (densities and concentrations) of common acid and base solutions. Nominal values are provided in cases where the manufacturer cites a range of concentrations and densities.
2. 2 This column contains specific gravity data. In the case of this table, specific gravity is the ratio of density of a substance to the density of pure water at the same conditions. Specific gravity is often cited on commercial labels.
3. 3 This solution is sometimes called “ammonium hydroxide,” although this term is not chemically accurate.

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13.1.5: Standard Thermodynamic Properties for Selected Substances

Standard Thermodynamic Properties for Selected Substances			
Substance	ΔH_f° (kJ mol^{-})	ΔG_f° (kJ mol^{-1})	S° (J K^{-1} mol^{-1})
aluminum			
Al(s)	0	0	28.3
Al(g)	324.4	285.7	164.54
Al_2O_3(s)	–1676	–1582	50.92
AlF_3(s)	–1510.4	–1425	66.5
$AlCl_3$(s)	–704.2	–628.8	110.67
$AlCl_3 \cdot 6H_2O$(s)	–2691.57	–2269.40	376.56
Al_2S_3(s)	–724.0	–492.4	116.9
$Al_2(SO_4)_3$(s)	–3445.06	–3506.61	239.32
antimony			
Sb(s)	0	0	45.69
Sb(g)	262.34	222.17	180.16
Sb_4O_6(s)	–1440.55	–1268.17	220.92
$SbCl_3$(g)	–313.8	–301.2	337.80
$SbCl_5$(g)	–394.34	–334.29	401.94
Sb_2S_3(s)	–174.89	–173.64	182.00
$SbCl_3$(s)	–382.17	–323.72	184.10
SbOCl(s)	–374.0	—	—
arsenic			
As(s)	0	0	35.1
As(g)	302.5	261.0	174.21
As_4(g)	143.9	92.4	314
As_4O_6(s)	–1313.94	–1152.52	214.22
As_2O_5(s)	–924.87	–782.41	105.44
$AsCl_3$(g)	–261.50	–248.95	327.06
As_2S_3(s)	–169.03	–168.62	163.59
AsH_3(g)	66.44	68.93	222.78
H_3AsO_4(s)	–906.3	—	—
barium			
Ba(s)	0	0	62.5
Ba(g)	180	146	170.24
BaO(s)	–548.0	–520.3	72.1

Standard Thermodynamic Properties for Selected Substances			
Substance	ΔH_f° (kJ mol^{-})	ΔG_f° (kJ mol^{-1})	S° (J K^{-1} mol^{-1})
$BaCl_2(s)$	–855.0	–806.7	123.7
$BaSO_4(s)$	–1473.2	–1362.3	132.2
beryllium			
$Be(s)$	0	0	9.50
$Be(g)$	324.3	286.6	136.27
$BeO(s)$	–609.4	–580.1	13.8
bismuth			
$Bi(s)$	0	0	56.74
$Bi(g)$	207.1	168.2	187.00
$Bi_2O_3(s)$	–573.88	–493.7	151.5
$BiCl_3(s)$	–379.07	–315.06	176.98
$Bi_2S_3(s)$	–143.1	–140.6	200.4
boron			
$B(s)$	0	0	5.86
$B(g)$	565.0	521.0	153.4
$B_2O_3(s)$	–1273.5	–1194.3	53.97
$B_2H_6(g)$	36.4	87.6	232.1
$H_3BO_3(s)$	–1094.33	–968.92	88.83
$BF_3(g)$	–1136.0	–1119.4	254.4
$BCl_3(g)$	–403.8	–388.7	290.1
$B_3N_3H_6(l)$	–540.99	–392.79	199.58
$HBO_2(s)$	–794.25	–723.41	37.66
bromine			
$Br_2(l)$	0	0	152.23
$Br_2(g)$	30.91	3.142	245.5
$Br(g)$	111.88	82.429	175.0
$BrF_3(g)$	–255.60	–229.45	292.42
$HBr(g)$	–36.3	–53.43	198.7
cadmium			
$Cd(s)$	0	0	51.76
$Cd(g)$	112.01	77.41	167.75
$CdO(s)$	–258.2	–228.4	54.8

Standard Thermodynamic Properties for Selected Substances			
Substance	ΔH_f° (kJ mol^{-})	ΔG_f° (kJ mol^{-1})	S° (J K^{-1} mol^{-1})
$CdCl_2(s)$	–391.5	–343.9	115.3
$CdSO_4(s)$	–933.3	–822.7	123.0
$CdS(s)$	–161.9	–156.5	64.9
calcium			
$Ca(s)$	0	0	41.6
$Ca(g)$	178.2	144.3	154.88
$CaO(s)$	–634.9	–603.3	38.1
$Ca(OH)_2(s)$	–985.2	–897.5	83.4
$CaSO_4(s)$	–1434.5	–1322.0	106.5
$CaSO_4 \cdot 2H_2O(s)$	–2022.63	–1797.45	194.14
$CaCO_3(s)$ (calcite)	–1220.0	–1081.4	110.0
$CaSO_3 \cdot H_2O(s)$	–1752.68	–1555.19	184.10
carbon			
$C(s)$ (graphite)	0	0	5.740
$C(s)$ (diamond)	1.89	2.90	2.38
$C(g)$	716.681	671.2	158.1
$CO(g)$	–110.52	–137.15	197.7
$CO_2(g)$	–393.51	–394.36	213.8
$CH_4(g)$	–74.6	–50.5	186.3
$CH_3OH(l)$	–239.2	–166.6	126.8
$CH_3OH(g)$	–201.0	–162.3	239.9
$CCl_4(l)$	–128.2	–62.5	214.4
$CCl_4(g)$	–95.7	–58.2	309.7
$CHCl_3(l)$	–134.1	–73.7	201.7
$CHCl_3(g)$	–103.14	–70.34	295.71
$CS_2(l)$	89.70	65.27	151.34
$CS_2(g)$	116.9	66.8	238.0
$C_2H_2(g)$	227.4	209.2	200.9
$C_2H_4(g)$	52.4	68.4	219.3
$C_2H_6(g)$	–84.0	–32.0	229.2
$CH_3CO_2H(l)$	–484.3	–389.9	159.8
$CH_3CO_2H(g)$	–434.84	–376.69	282.50

Standard Thermodynamic Properties for Selected Substances			
Substance	ΔH_f° (kJ mol^{-})	ΔG_f° (kJ mol^{-1})	S° (J K^{-1} mol^{-1})
$C_2H_5OH(l)$	–277.6	–174.8	160.7
$C_2H_5OH(g)$	–234.8	–167.9	281.6
$C_3H_8(g)$	–103.8	–23.4	270.3
$C_6H_6(g)$	82.927	129.66	269.2
$C_6H_6(l)$	49.1	124.50	173.4
$CH_2Cl_2(l)$	–124.2	–63.2	177.8
$CH_2Cl_2(g)$	–95.4	–65.90	270.2
$CH_3Cl(g)$	–81.9	–60.2	234.6
$C_2H_5Cl(l)$	–136.52	–59.31	190.79
$C_2H_5Cl(g)$	–112.17	–60.39	276.00
$C_2N_2(g)$	308.98	297.36	241.90
$HCN(l)$	108.9	125.0	112.8
$HCN(g)$	135.5	124.7	201.8
chlorine			
$Cl_2(g)$	0	0	223.1
$Cl(g)$	121.3	105.70	165.2
$ClF(g)$	–54.48	–55.94	217.78
$ClF_3(g)$	–158.99	–118.83	281.50
$Cl_2O(g)$	80.3	97.9	266.2
$Cl_2O_7(l)$	238.1	—	—
$Cl_2O_7(g)$	272.0	—	—
$HCl(g)$	–92.307	–95.299	186.9
$HClO_4(l)$	–40.58	—	—
chromium			
$Cr(s)$	0	0	23.77
$Cr(g)$	396.6	351.8	174.50
$Cr_2O_3(s)$	–1139.7	–1058.1	81.2
$CrO_3(s)$	–589.5	—	—
$(NH_4)_2Cr_2O_7(s)$	–1806.7	—	—
cobalt			
$Co(s)$	0	0	30.0
$CoO(s)$	–237.9	–214.2	52.97

Standard Thermodynamic Properties for Selected Substances			
Substance	ΔH_f° (kJ mol^{-1})	ΔG_f° (kJ mol^{-1})	S° (J K^{-1} mol^{-1})
$Co_3O_4(s)$	–910.02	–794.98	114.22
$Co(NO_3)_2(s)$	–420.5	—	—
copper			
$Cu(s)$	0	0	33.15
$Cu(g)$	338.32	298.58	166.38
$CuO(s)$	–157.3	–129.7	42.63
$Cu_2O(s)$	–168.6	–146.0	93.14
$CuS(s)$	–53.1	–53.6	66.5
$Cu_2S(s)$	–79.5	–86.2	120.9
$CuSO_4(s)$	–771.36	–662.2	109.2
$Cu(NO_3)_2(s)$	–302.9	—	—
fluorine			
$F_2(g)$	0	0	202.8
$F(g)$	79.4	62.3	158.8
$F_2O(g)$	24.7	41.9	247.43
$HF(g)$	–273.3	–275.4	173.8
hydrogen			
$H_2(g)$	0	0	130.7
$H(g)$	217.97	203.26	114.7
$H_2O(l)$	–285.83	–237.1	70.0
$H_2O(g)$	–241.82	–228.59	188.8
$H_2O_2(l)$	–187.78	–120.35	109.6
$H_2O_2(g)$	–136.3	–105.6	232.7
$HF(g)$	–273.3	–275.4	173.8
$HCl(g)$	–92.307	–95.299	186.9
$HBr(g)$	–36.3	–53.43	198.7
$HI(g)$	26.48	1.70	206.59
$H_2S(g)$	–20.6	–33.4	205.8
$H_2Se(g)$	29.7	15.9	219.0
iodine			
$I_2(s)$	0	0	116.14
$I_2(g)$	62.438	19.3	260.7

Standard Thermodynamic Properties for Selected Substances			
Substance	ΔH_f° (kJ mol^{-})	ΔG_f° (kJ mol^{-1})	S° (J K^{-1} mol^{-1})
I(g)	106.84	70.2	180.8
IF(g)	95.65	–118.49	236.06
ICl(g)	17.78	–5.44	247.44
IBr(g)	40.84	3.72	258.66
$IF_7(g)$	–943.91	–818.39	346.44
HI(g)	26.48	1.70	206.59
iron			
Fe(s)	0	0	27.3
Fe(g)	416.3	370.7	180.5
$Fe_2O_3(s)$	–824.2	–742.2	87.40
$Fe_3O_4(s)$	–1118.4	–1015.4	146.4
$Fe(CO)_5(l)$	–774.04	–705.42	338.07
$Fe(CO)_5(g)$	–733.87	–697.26	445.18
$FeCl_2(s)$	–341.79	–302.30	117.95
$FeCl_3(s)$	–399.49	–334.00	142.3
FeO(s)	–272.0	–255.2	60.75
$Fe(OH)_2(s)$	–569.0	–486.5	88.
$Fe(OH)_3(s)$	–823.0	–696.5	106.7
FeS(s)	–100.0	–100.4	60.29
$Fe_3C(s)$	25.10	20.08	104.60
lead			
Pb(s)	0	0	64.81
Pb(g)	195.2	162.	175.4
PbO(s) (yellow)	–217.32	–187.89	68.70
PbO(s) (red)	–218.99	–188.93	66.5
$Pb(OH)_2(s)$	–515.9	—	—
PbS(s)	–100.4	–98.7	91.2
$Pb(NO_3)_2(s)$	–451.9	—	—
$PbO_2(s)$	–277.4	–217.3	68.6
$PbCl_2(s)$	–359.4	–314.1	136.0
lithium			
Li(s)	0	0	29.1

Standard Thermodynamic Properties for Selected Substances			
Substance	ΔH_f° (kJ mol^{-})	ΔG_f° (kJ mol^{-1})	S° (J K^{-1} mol^{-1})
Li(g)	159.3	126.6	138.8
LiH(s)	–90.5	–68.3	20.0
Li(OH)(s)	–487.5	–441.5	42.8
LiF(s)	–616.0	–587.5	35.7
Li_2CO_3(s)	–1216.04	–1132.19	90.17
manganese			
Mn(s)	0	0	32.0
Mn(g)	280.7	238.5	173.7
MnO(s)	–385.2	–362.9	59.71
MnO_2(s)	–520.03	–465.1	53.05
Mn_2O_3(s)	–958.97	–881.15	110.46
Mn_3O_4(s)	–1378.83	–1283.23	155.64
mercury			
Hg(l)	0	0	75.9
Hg(g)	61.4	31.8	175.0
HgO(s) (red)	–90.83	–58.5	70.29
HgO(s) (yellow)	–90.46	–58.43	71.13
$HgCl_2$(s)	–224.3	–178.6	146.0
Hg_2Cl_2(s)	–265.4	–210.7	191.6
HgS(s) (red)	–58.16	–50.6	82.4
HgS(s) (black)	–53.56	–47.70	88.28
$HgSO_4$(s)	–707.51	–594.13	0.00
nitrogen			
N_2(g)	0	0	191.6
N(g)	472.704	455.5	153.3
NO(g)	90.25	87.6	210.8
NO_2(g)	33.2	51.30	240.1
N_2O(g)	81.6	103.7	220.0
N_2O_3(g)	83.72	139.41	312.17
N_2O_4(g)	11.1	99.8	304.4
N_2O_5(g)	11.3	115.1	355.7
NH_3(g)	–45.9	–16.5	192.8

Standard Thermodynamic Properties for Selected Substances			
Substance	ΔH_f° (kJ mol^{-1})	ΔG_f° (kJ mol^{-1})	S° (J K^{-1} mol^{-1})
$N_2H_4(l)$	50.63	149.43	121.21
$N_2H_4(g)$	95.4	159.4	238.5
$NH_4NO_3(s)$	–365.56	–183.87	151.08
$NH_4Cl(s)$	–314.43	–202.87	94.6
$NH_4Br(s)$	–270.8	–175.2	113.0
$NH_4I(s)$	–201.4	–112.5	117.0
$NH_4NO_2(s)$	–256.5	—	—
$HNO_3(l)$	–174.1	–80.7	155.6
$HNO_3(g)$	–133.9	–73.5	266.9
oxygen			
$O_2(g)$	0	0	205.2
$O(g)$	249.17	231.7	161.1
$O_3(g)$	142.7	163.2	238.9
phosphorus			
$P_4(s)$	0	0	164.4
$P_4(g)$	58.91	24.4	280.0
$P(g)$	314.64	278.25	163.19
$PH_3(g)$	5.4	13.5	210.2
$PCl_3(g)$	–287.0	–267.8	311.78
$PCl_5(g)$	–374.9	–305.0	364.4
$P_4O_6(s)$	–1640.1	—	—
$P_4O_{10}(s)$	–2984.0	–2697.0	228.86
$HPO_3(s)$	–948.5	—	—
$H_3PO_2(s)$	–604.6	—	—
$H_3PO_3(s)$	–964.4	—	—
$H_3PO_4(s)$	–1279.0	–1119.1	110.50
$H_3PO_4(l)$	–1266.9	–1124.3	110.5
$H_4P_2O_7(s)$	–2241.0	—	—
$POCl_3(l)$	–597.1	–520.8	222.5
$POCl_3(g)$	–558.5	–512.9	325.5
potassium			
$K(s)$	0	0	64.7

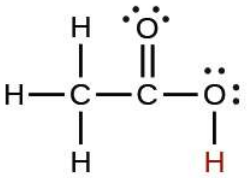
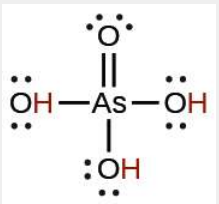
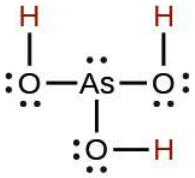
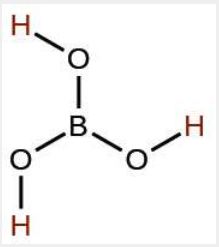
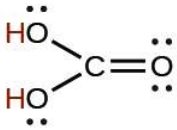
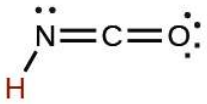
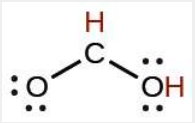
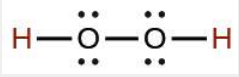
Standard Thermodynamic Properties for Selected Substances			
Substance	ΔH_f° (kJ mol^{-})	ΔG_f° (kJ mol^{-1})	S° (J K^{-1} mol^{-1})
K(*g*)	89.0	60.5	160.3
KF(*s*)	–576.27	–537.75	66.57
KCl(*s*)	–436.5	–408.5	82.6
silicon			
Si(*s*)	0	0	18.8
Si(*g*)	450.0	405.5	168.0
SiO_2(*s*)	–910.7	–856.3	41.5
SiH_4(*g*)	34.3	56.9	204.6
H_2SiO_3(*s*)	–1188.67	–1092.44	133.89
H_4SiO_4(*s*)	–1481.14	–1333.02	192.46
SiF_4(*g*)	–1615.0	–1572.8	282.8
$SiCl_4$(*l*)	–687.0	–619.8	239.7
$SiCl_4$(*g*)	–662.75	–622.58	330.62
SiC(*s, beta cubic*)	–73.22	–70.71	16.61
SiC(*s, alpha hexagonal*)	–71.55	–69.04	16.48
silver			
Ag(*s*)	0	0	42.55
Ag(*g*)	284.9	246.0	172.89
Ag_2O(*s*)	–31.05	–11.20	121.3
AgCl(*s*)	–127.0	–109.8	96.3
Ag_2S(*s*)	–32.6	–40.7	144.0
sodium			
Na(*s*)	0	0	51.3
Na(*g*)	107.5	77.0	153.7
Na_2O(*s*)	–414.2	–375.5	75.1
NaCl(*s*)	–411.2	–384.1	72.1
sulfur			
S_8(*s*) (rhombic)	0	0	256.8
S(*g*)	278.81	238.25	167.82
SO_2(*g*)	–296.83	–300.1	248.2
SO_3(*g*)	–395.72	–371.06	256.76
H_2S(*g*)	–20.6	–33.4	205.8

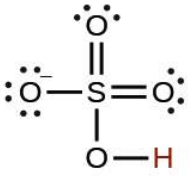
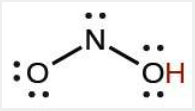
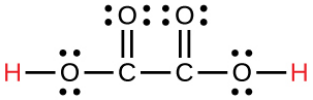
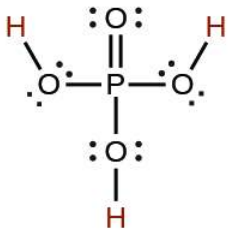
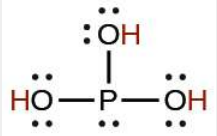
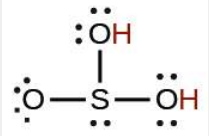
Standard Thermodynamic Properties for Selected Substances			
Substance	ΔH_f° (kJ mol^{-})	ΔG_f° (kJ mol^{-1})	S° (J K^{-1} mol^{-1})
$H_2SO_4(l)$	–813.989	690.00	156.90
$H_2S_2O_7(s)$	–1273.6	—	—
$SF_4(g)$	–728.43	–684.84	291.12
$SF_6(g)$	–1220.5	–1116.5	291.5
$SCl_2(l)$	–50	—	—
$SCl_2(g)$	–19.7	—	—
$S_2Cl_2(l)$	–59.4	—	—
$S_2Cl_2(g)$	–19.50	–29.25	319.45
$SOCl_2(g)$	–212.55	–198.32	309.66
$SOCl_2(l)$	–245.6	—	—
$SO_2Cl_2(l)$	–394.1	—	—
SO_2Cl_2(g)	–354.80	–310.45	311.83
tin			
Sn(s)	0	0	51.2
Sn(g)	301.2	266.2	168.5
SnO(s)	–285.8	–256.9	56.5
$SnO_2(s)$	–577.6	–515.8	49.0
$SnCl_4(l)$	–511.3	–440.1	258.6
$SnCl_4(g)$	–471.5	–432.2	365.8
titanium			
Ti(s)	0	0	30.7
Ti(g)	473.0	428.4	180.3
$TiO_2(s)$	–944.0	–888.8	50.6
$TiCl_4(l)$	–804.2	–737.2	252.4
$TiCl_4(g)$	–763.2	–726.3	353.2
tungsten			
W(s)	0	0	32.6
W(g)	849.4	807.1	174.0
$WO_3(s)$	–842.9	–764.0	75.9
zinc			
Zn(s)	0	0	41.6
Zn(g)	130.73	95.14	160.98

Standard Thermodynamic Properties for Selected Substances			
Substance	ΔH_f° (kJ mol^{-})	ΔG_f° (kJ mol^{-1})	S_{298}° (J K^{-1} mol^{-1})
ZnO(*s*)	–350.5	–320.5	43.7
$ZnCl_2$(*s*)	–415.1	–369.43	111.5
ZnS(*s*)	–206.0	–201.3	57.7
$ZnSO_4$(*s*)	–982.8	–871.5	110.5
$ZnCO_3$(*s*)	–812.78	–731.57	82.42
complexes			
$[Co(NH_3)_4(NO_2)_2]NO_3$, *cis*	–898.7	—	—
$[Co(NH_3)_4(NO_2)_2]NO_3$, *trans*	–896.2	—	—
$NH_4[Co(NH_3)_2(NO_2)_4]$	–837.6	—	—
$[Co(NH_3)_6][Co(NH_3)_2(NO_2)_4]_3$	–2733.0	—	—
$[Co(NH_3)_4Cl_2]Cl$, *cis*	–874.9	—	—
$[Co(NH_3)_4Cl_2]Cl$, *trans*	–877.4	—	—
$[Co(en)_2(NO_2)_2]NO_3$, *cis*	–689.5	—	—
$[Co(en)_2Cl_2]Cl$, *cis*	–681.2	—	—
$[Co(en)_2Cl_2]Cl$, *trans*	–677.4	—	—
$[Co(en)_3](ClO_4)_3$	–762.7	—	—
$[Co(en)_3]Br_2$	–595.8	—	—
$[Co(en)_3]I_2$	–475.3	—	—
$[Co(en)_3]I_3$	–519.2	—	—
$[Co(NH_3)_6](ClO_4)_3$	–1034.7	–221.1	615
$[Co(NH_3)_5NO_2](NO_3)_2$	–1088.7	–412.9	331
$[Co(NH_3)_6](NO_3)_3$	–1282.0	–524.5	448
$[Co(NH_3)_5Cl]Cl_2$	–1017.1	–582.5	366.1
$[Pt(NH_3)_4]Cl_2$	–725.5	—	—
$[Ni(NH_3)_6]Cl_2$	–994.1	—	—
$[Ni(NH_3)_6]Br_2$	–923.8	—	—
$[Ni(NH_3)_6]I_2$	–808.3	—	—

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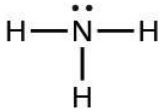
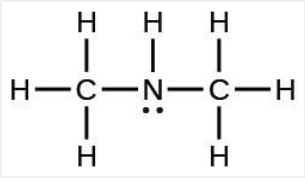
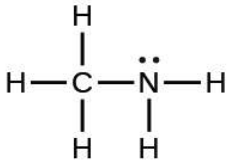
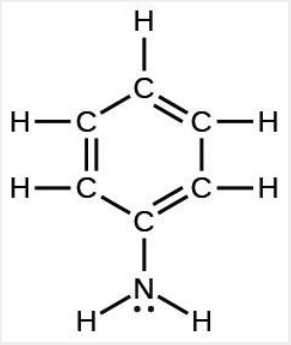
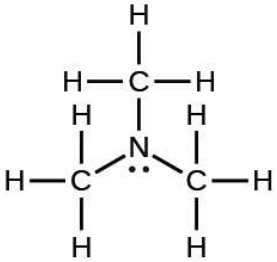
13.2.6: Ionization Constants of Weak Acids

Ionization Constants of Weak Acids			
Acid	Formula	K_a at 25 °C	Lewis Structure
acetic	$CH_3CO_2\mathbf{H}$	1.8×10^{-5}	
arsenic	$\mathbf{H_3}AsO_4$	5.5×10^{-3}	
	$H_2AsO_4^-$	1.7×10^{-7}	
	$HAsO_4^{2-}$	5.1×10^{-12}	
arsenous	$\mathbf{H_3}AsO_3$	5.1×10^{-10}	
boric	$\mathbf{H_3}BO_3$	5.4×10^{-10}	
carbonic	$\mathbf{H_2}CO_3$	4.3×10^{-7}	
	HCO_3^-	5.6×10^{-11}	
cyanic	**H**CNO	2×10^{-4}	
formic	$HCO_2\mathbf{H}$	1.8×10^{-4}	
hydrazoic	$\mathbf{H}N_3$	2.5×10^{-5}	
hydrocyanic	**H**CN	4.9×10^{-10}	
hydrofluoric	**H**F	3.5×10^{-4}	
hydrogen peroxide	$\mathbf{H_2}O_2$	2.4×10^{-12}	

Ionization Constants of Weak Acids			
Acid	Formula	K_a at 25 °C	Lewis Structure
hydrogen selenide	H_2Se	1.29×10^{-4}	
	HSe^-	1×10^{-12}	
hydrogen sulfate ion	HSO_4^-	1.2×10^{-2}	
hydrogen sulfide	H_2S	8.9×10^{-8}	
	HS^-	1.0×10^{-19}	
hydrogen telluride	H_2Te	2.3×10^{-3}	
	HTe^-	1.6×10^{-11}	
hypobromous	HBrO	2.8×10^{-9}	
hypochlorous	HClO	2.9×10^{-8}	
nitrous	HNO_2	4.6×10^{-4}	
oxalic	$H_2C_2O_4$	6.0×10^{-2}	
	$HC_2O_4^-$	6.1×10^{-5}	
phosphoric	H_3PO_4	7.5×10^{-3}	
	$H_2PO_4^-$	6.2×10^{-8}	
	HPO_4^{2-}	4.2×10^{-13}	
phosphorous	H_3PO_3	5×10^{-2}	
	$H_2PO_3^-$	2.0×10^{-7}	
sulfurous	H_2SO_3	1.6×10^{-2}	
	HSO_3^-	6.4×10^{-8}	

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13.2.7: Ionization Constants of Weak Bases

Ionization Constants of Weak Bases		
Base	Lewis Structure	K_b at 25 °C
ammonia		1.8×10^{-5}
dimethylamine		5.9×10^{-4}
methylamine		4.4×10^{-4}
phenylamine (aniline)		4.3×10^{-10}
trimethylamine		6.3×10^{-5}

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13.1.8: Solubility Products

Solubility Products	
Substance	K_{sp} at 25 °C
aluminum	
$Al(OH)_3$	2×10^{-32}
barium	
$BaCO_3$	1.6×10^{-9}
$BaC_2O_4 \cdot 2H_2O$	1.1×10^{-7}
$BaSO_4$	2.3×10^{-8}
$BaCrO_4$	8.5×10^{-11}
BaF_2	2.4×10^{-5}
$Ba(OH)_2 \cdot 8H_2O$	5.0×10^{-3}
$Ba_3(PO_4)_2$	6×10^{-39}
$Ba_3(AsO_4)_2$	1.1×10^{-13}
bismuth	
BiO(OH)	4×10^{-10}
BiOCl	1.8×10^{-31}
Bi_2S_3	1×10^{-97}
cadmium	
$Cd(OH)_2$	5.9×10^{-15}
CdS	1.0×10^{-28}
$CdCO_3$	5.2×10^{-12}
calcium	
$Ca(OH)_2$	1.3×10^{-6}
$CaCO_3$	8.7×10^{-9}
$CaSO4 \cdot 2H_2O$	6.1×10^{-5}
$CaC_2O_4 \cdot H_2O$	1.96×10^{-8}
$Ca_3(PO_4)_2$	1.3×10^{-32}
$CaHPO_4$	7×10^{-7}
CaF_2	4.0×10^{-11}
chromium	
$Cr(OH)_3$	6.7×10^{-31}
cobalt	
$Co(OH)_2$	2.5×10^{-16}
$CoS(\alpha)$	5×10^{-22}

Solubility Products	
Substance	K_{sp} at 25 °C
$CoS(\beta)$	3×10^{-26}
$CoCO_3$	1.4×10^{-13}
$Co(OH)_3$	2.5×10^{-43}
copper	
CuCl	1.2×10^{-6}
CuBr	6.27×10^{-9}
CuI	1.27×10^{-12}
CuSCN	1.6×10^{-11}
Cu_2S	2.5×10^{-48}
$Cu(OH)_2$	2.2×10^{-20}
CuS	8.5×10^{-45}
$CuCO_3$	2.5×10^{-10}
iron	
$Fe(OH)_2$	1.8×10^{-15}
$FeCO_3$	2.1×10^{-11}
FeS	3.7×10^{-19}
$Fe(OH)_3$	4×10^{-38}
lead	
$Pb(OH)_2$	1.2×10^{-15}
PbF_2	4×10^{-8}
$PbCl_2$	1.6×10^{-5}
$PbBr_2$	4.6×10^{-6}
PbI_2	1.4×10^{-8}
$PbCO_3$	1.5×10^{-15}
PbS	7×10^{-29}
$PbCrO_4$	2×10^{-16}
$PbSO_4$	1.3×10^{-8}
$Pb_3(PO_4)_2$	1×10^{-54}
magnesium	
$Mg(OH)_2$	8.9×10^{-12}
$MgCO_3 \cdot 3H_2O$	*ca* 1×10^{-5}
$MgNH_4PO_4$	3×10^{-13}

Solubility Products	
Substance	K_{sp} at 25 °C
MgF_2	6.4×10^{-9}
MgC_2O_4	7×10^{-7}
manganese	
$Mn(OH)_2$	2×10^{-13}
$MnCO_3$	8.8×10^{-11}
MnS	2.3×10^{-13}
mercury	
$Hg_2O{\cdot}H_2O$	3.6×10^{-26}
Hg_2Cl_2	1.1×10^{-18}
Hg_2Br_2	1.3×10^{-22}
Hg_2I_2	4.5×10^{-29}
Hg_2CO_3	9×10^{-15}
Hg_2SO_4	7.4×10^{-7}
Hg_2S	1.0×10^{-47}
Hg_2CrO_4	2×10^{-9}
HgS	1.6×10^{-54}
nickel	
$Ni(OH)_2$	1.6×10^{-16}
$NiCO_3$	1.4×10^{-7}
NiS(α)	4×10^{-20}
NiS(β)	1.3×10^{-25}
potassium	
$KClO_4$	1.05×10^{-2}
K_2PtCl_6	7.48×10^{-6}
$KHC_4H_4O_6$	3×10^{-4}
silver	
$\frac{1}{2}Ag_2O(Ag^+ + OH^-)$	2×10^{-8}
AgCl	1.6×10^{-10}
AgBr	5.0×10^{-13}
AgI	1.5×10^{-16}
AgCN	1.2×10^{-16}
AgSCN	1.0×10^{-12}

Solubility Products	
Substance	K_{sp} at 25 °C
Ag_2S	1.6×10^{-49}
Ag_2CO_3	8.1×10^{-12}
Ag_2CrO_4	9.0×10^{-12}
$Ag_4Fe(CN)_6$	1.55×10^{-41}
Ag_2SO_4	1.2×10^{-5}
Ag_3PO_4	1.8×10^{-18}
strontium	
$Sr(OH)_2 \cdot 8H_2O$	3.2×10^{-4}
$SrCO_3$	7×10^{-10}
$SrCrO_4$	3.6×10^{-5}
$SrSO_4$	3.2×10^{-7}
$SrC_2O_4 \cdot H_2O$	4×10^{-7}
thallium	
TlCl	1.7×10^{-4}
TlSCN	1.6×10^{-4}
Tl_2S	6×10^{-22}
$Tl(OH)_3$	6.3×10^{-46}
tin	
$Sn(OH)_2$	3×10^{-27}
SnS	1×10^{-26}
$Sn(OH)_4$	1.0×10^{-57}
zinc	
$ZnCO_3$	2×10^{-10}

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13.1.9: Formation Constants for Complex Ions

Formation Constants for Complex Ions	
Equilibrium	K_f
$Al^{3+} + 6\,F^- \rightleftharpoons [AlF_6]^{3-}$	7×10^{19}
$Cd^{2+} + 4\,NH_3 \rightleftharpoons [Cd(NH_3)_4]^{2+}$	1.3×10^{7}
$Cd^{2+} + 4\,CN^- \rightleftharpoons [Cd(CN)_4]^{2-}$	3×10^{18}
$Co^{2+} + 6\,NH_3 \rightleftharpoons [Co(NH_3)_6]^{2+}$	1.3×10^{5}
$Co^{3+} + 6\,NH_3 \rightleftharpoons [Co(NH_3)_6]^{3+}$	2.3×10^{33}
$Cu^{+} + 2\,CN \rightleftharpoons [Cu(CN)_2]^-$	1.0×10^{16}
$Cu^{2+} + 4\,NH_3 \overset{Cu(NH_3)_4}{\rightleftharpoons} {}^{2+}$	1.7×10^{13}
$Fe^{2+} + 6\,CN^- \overset{Fe(CN)_6}{\rightleftharpoons} {}^{4-}$	1.5×10^{35}
$Fe^{3+} + 6\,CN^- \overset{Fe(CN)_6}{\rightleftharpoons} {}^{3-}$	2×10^{43}
$Fe^{3+} + 6\,SCN^- \overset{Fe(SCN)_6}{\rightleftharpoons} {}^{3-}$	3.2×10^{3}
$Hg^{2+} + 4\,Cl^- \overset{HgCl_4}{\rightleftharpoons} {}^{2-}$	1.1×10^{16}
$Ni^{2+} + 6\,NH_3 \overset{Ni(NH_3)_6}{\rightleftharpoons} {}^{2+}$	2.0×10^{8}
$Ag^{+} + 2\,Cl^- \overset{AgCl_2}{\rightleftharpoons} {}^{-}$	1.8×10^{5}
$Ag^{+} + 2\,CN^- \overset{Ag(CN)_2}{\rightleftharpoons} {}^{-}$	1×10^{21}
$Ag^{+} + 2\,NH_3 \overset{Ag(NH_3)_2}{\rightleftharpoons} {}^{+}$	1.7×10^{7}
$Zn^{2+} + 4\,CN^- \overset{Zn(CN)_4}{\rightleftharpoons} {}^{2-}$	2.1×10^{19}
$Zn^{2+} + 4\,OH^- \overset{Zn(OH)_4}{\rightleftharpoons} {}^{2-}$	2×10^{15}
$Fe^{3+} + SCN^- \overset{Fe(SCN)}{\rightleftharpoons} {}^{2+}$	8.9×10^{2}
$Ag^{+} + 4\,SCN^- \overset{Ag(SCN)_4}{\rightleftharpoons} {}^{3-}$	1.2×10^{10}
$Pb^{2+} + 4\,I^- \overset{PbI_4}{\rightleftharpoons} {}^{2-}$	3.0×10^{4}
$Pt^{2+} + 4\,Cl^- \overset{PtCl_4}{\rightleftharpoons} {}^{2-}$	1×10^{16}
$Cu^{2+} + 4\,CN \overset{Cu(CN)_4}{\rightleftharpoons} {}^{2-}$	1.0×10^{25}
$Co^{2+} + 4\,SCN^- \overset{Co(SCN)_4}{\rightleftharpoons} {}^{2-}$	1×10^{3}

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13.2.10: Standard Electrode (Half-Cell) Potentials

Standard Electrode (Half-Cell) Potentials	
Half-Reaction	$E°$ (V)
$Ag^+ + e^- \longrightarrow Ag$	+0.7996
$AgCl + e^- \longrightarrow Ag + Cl^-$	+0.22233
$[Ag(CN)_2]^- + e^- \longrightarrow Ag + 2\,CN^-$	−0.31
$Ag_2CrO_4 + 2\,e^- \longrightarrow 2\,Ag + CrO_4^{2-}$	+0.45
$[Ag(NH_3)_2]^+ + e^- \longrightarrow Ag + 2\,NH_3$	+0.373
$[Ag(S_2O_3)_2]^{3+} + e^- \longrightarrow Ag + 2\,S_2O_3^{2-}$	+0.017
$[AlF_6]^{3-} + 3\,e^- \longrightarrow Al + 6\,F^-$	−2.07
$Al^{3+} + 3\,e^- \longrightarrow Al$	−1.662
$Am^{3+} + 3\,e^- \longrightarrow Am$	−2.048
$Au^{3+} + 3\,e^- \longrightarrow Au$	+1.498
$Au^+ + e^- \longrightarrow Au$	+1.692
$Ba^{2+} + 2\,e^- \longrightarrow Ba$	−2.912
$Be^{2+} + 2\,e^- \longrightarrow Be$	−1.847
$Br_2(aq) + 2\,e^- \longrightarrow 2\,Br^-$	+1.0873
$Ca^{2+} + 2\,e^- \longrightarrow Ca$	−2.868
$Ce^3 + 3\,e^- \longrightarrow Ce$	−2.483
$Ce^{4+} + e^- \longrightarrow Ce^{3+}$	+1.61
$Cd^{2+} + 2\,e^- \longrightarrow Cd$	−0.4030
$[Cd(CN)_4]^{2-} + 2\,e^- \longrightarrow Cd + 4\,CN^-$	−1.09
$[Cd(NH_3)_4]^{2+} + 2\,e^- \longrightarrow Cd + 4\,NH_3$	−0.61
$CdS + 2\,e^- \longrightarrow Cd + S^{2-}$	−1.17
$Cl_2 + 2\,e^- \longrightarrow 2\,Cl^-$	+1.35827
$ClO_4^- + H_2O + 2\,e^- \longrightarrow ClO_3^- + 2\,OH^-$	+0.36
$ClO_3^- + H_2O + 2\,e^- \longrightarrow ClO_2^- + 2\,OH^-$	+0.33
$ClO_2^- + H_2O + 2\,e^- \longrightarrow ClO^- + 2\,OH^-$	+0.66
$ClO^- + H_2O + 2\,e^- \longrightarrow Cl^- + 2\,OH^-$	+0.89
$ClO_4^- + 2\,H_3O^+ + 2\,e^- \longrightarrow ClO_3^- + 3\,H_2O$	+1.189
$ClO_3^- + 3\,H_3O^+ + 2\,e^- \longrightarrow HClO_2 + 4\,H_2O$	+1.21
$HClO + H_3O^+ + 2\,e^- \longrightarrow Cl^- + 2\,H_2O$	+1.482
$HClO + H_3O^+ + e^- \longrightarrow \frac{1}{2}Cl_2 + 2\,H_2O$	+1.611
$HClO_2 + 2\,H_3O^+ + 2\,e^- \longrightarrow HClO + 3\,H_2O$	+1.628
$Co^{3+} + e^- \longrightarrow Co^{2+}\ \ (2\,mol\,//\,H_2SO_4)$	+1.83

Standard Electrode (Half-Cell) Potentials	
Half-Reaction	$E°$ (V)
$Co^{2+} + 2\,e^- \longrightarrow Co$	−0.28
$[Co(NH_3)_6]^{3+} + e^- \xrightarrow{Co(NH_3)_6} {}^{2+}$	+0.1
$Co(OH)_3 + e^- \longrightarrow Co(OH)_2 + OH^-$	+0.17
$Cr^3 + 3\,e^- \longrightarrow Cr$	−0.744
$Cr^{3+} + e^- \longrightarrow Cr^{2+}$	−0.407
$Cr^{2+} + 2\,e^- \longrightarrow Cr$	−0.913
$[Cu(CN)_2]^- + e^- \longrightarrow Cu + 2\,CN^-$	−0.43
$CrO_4{}^{2-} + 4\,H_2O + 3\,e^- \longrightarrow Cr(OH)_3 + 5\,OH^-$	−0.13
$Cr_2O_7{}^{2-} + 14\,H_3O^+ + 6\,e^- \longrightarrow 2\,Cr^{3+} + 21\,H_2O$	+1.232
$[Cr(OH)_4]^- + 3\,e^- \longrightarrow Cr + 4\,OH^-$	−1.2
$Cr(OH)_3 + 3\,e^- \longrightarrow Cr + 3\,OH^-$	−1.48
$Cu^{2+} + e^- \longrightarrow Cu^+$	+0.153
$Cu^{2+} + 2\,e^- \longrightarrow Cu$	+0.34
$Cu^+ + e^- \longrightarrow Cu$	+0.521
$F_2 + 2\,e^- \longrightarrow 2\,F^-$	+2.866
$Fe^{2+} + 2\,e^- \longrightarrow Fe$	−0.447
$Fe^{3+} + e^- \longrightarrow Fe^{2+}$	+0.771
$[Fe(CN)_6]^{3-} + e^- \xrightarrow{Fe(CN)_6} {}^{4-}$	+0.36
$Fe(OH)_2 + 2\,e^- \longrightarrow Fe + 2\,OH^-$	−0.88
$FeS + 2\,e^- \longrightarrow Fe + S^{2-}$	−1.01
$Ga^{3+} + 3\,e^- \longrightarrow Ga$	−0.549
$Gd^{3+} + 3\,e^- \longrightarrow Gd$	−2.279
$\frac{1}{2}H_2 + e^- \longrightarrow H^-$	−2.23
$2\,H_2O + 2\,e^- \longrightarrow H_2 + 2\,OH^-$	−0.8277
$H_2O_2 + 2\,H_3O^+ + 2\,e^- \longrightarrow 4\,H_2O$	+1.776
$2\,H_3O^+ + 2\,e^- \longrightarrow H_2 + 2\,H_2O$	0.00
$HO_2{}^- + H_2O + 2\,e^- \longrightarrow 3\,OH^-$	+0.878
$Hf^{4+} + 4\,e^- \longrightarrow Hf$	−1.55
$Hg^{2+} + 2\,e^- \longrightarrow Hg$	+0.851
$2\,Hg^{2+} + 2\,e^- \longrightarrow Hg_2{}^{2+}$	+0.92
$Hg_2{}^{2+} + 2\,e^- \longrightarrow 2\,Hg$	+0.7973

Standard Electrode (Half-Cell) Potentials	
Half-Reaction	$E°$ (V)
$[HgBr_4]^{2-} + 2e^- \longrightarrow Hg + 4Br^-$	+0.21
$Hg_2Cl_2 + 2e^- \longrightarrow 2Hg + 2Cl^-$	+0.26808
$[Hg(CN)_4]^{2-} + 2e^- \longrightarrow Hg + 4CN^-$	−0.37
$[HgI_4]^{2-} + 2e^- \longrightarrow Hg + 4I^-$	−0.04
$HgS + 2e^- \longrightarrow Hg + S^{2-}$	−0.70
$I_2 + 2e^- \longrightarrow 2I^-$	+0.5355
$In^{3+} + 3e^- \longrightarrow In$	−0.3382
$K^+ + e^- \longrightarrow K$	−2.931
$La^{3+} + 3e^- \longrightarrow La$	−2.52
$Li^+ + e^- \longrightarrow Li$	−3.04
$Lu^{3+} + 3e^- \longrightarrow Lu$	−2.28
$Mg^{2+} + 2e^- \longrightarrow Mg$	−2.372
$Mn^{2+} + 2e^- \longrightarrow Mn$	−1.185
$MnO_2 + 2H_2O + 2e^- \longrightarrow Mn(OH)_2 + 2OH^-$	−0.05
$MnO_4^- + 2H_2O + 3e^- \longrightarrow MnO_2 + 4OH^-$	+0.558
$MnO_2 + 4H^+ + 2e^- \longrightarrow Mn^{2+} + 2H_2O$	+1.23
$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O$	+1.507
$Na^+ + e^- \longrightarrow Na$	−2.71
$Nd^{3+} + 3e^- \longrightarrow Nd$	−2.323
$Ni^{2+} + 2e^- \longrightarrow Ni$	−0.257
$[Ni(NH_3)_6]^{2+} + 2e^- \longrightarrow Ni + 6NH_3$	−0.49
$NiO_2 + 4H^+ + 2e^- \longrightarrow Ni^{2+} + 2H_2O$	+1.593
$NiO_2 + 2H_2O + 2e^- \longrightarrow Ni(OH)_2 + 2OH^-$	+0.49
$NiS + 2e^- \longrightarrow Ni + S^{2-}$	+0.76
$NO_3^- + 4H^+ + 3e^- \longrightarrow NO + 2H_2O$	+0.957
$NO_3^- + 3H^+ + 2e^- \longrightarrow HNO_2 + H_2O$	+0.92
$NO_3^- + H_2O + 2e^- \longrightarrow NO_2^- + 2OH^-$	+0.10
$Np^{3+} + 3e^- \longrightarrow Np$	−1.856
$O_2 + 2H_2O + 4e^- \longrightarrow 4OH^-$	+0.401
$O_2 + 2H^+ + 2e^- \longrightarrow H_2O_2$	+0.695
$O_2 + 4H^+ + 4e^- \longrightarrow 2H_2O$	+1.229
$Pb^{2+} + 2e^- \longrightarrow Pb$	−0.1262

Standard Electrode (Half-Cell) Potentials	
Half-Reaction	$E°$ (V)
$PbO_2 + SO_4^{2-} + 4H^+ + 2e^- \longrightarrow PbSO_4 + 2H_2O$	+1.69
$PbS + 2e^- \longrightarrow Pb + S^{2-}$	−0.95
$PbSO_4 + 2e^- \longrightarrow Pb + SO_4^{2-}$	−0.3505
$Pd^{2+} + 2e^- \longrightarrow Pd$	+0.987
$[PdCl_4]^{2-} + 2e^- \longrightarrow Pd + 4Cl^-$	+0.591
$Pt^{2+} + 2e^- \longrightarrow Pt$	+1.20
$[PtBr_4]^{2-} + 2e^- \longrightarrow Pt + 4Br^-$	+0.58
$[PtCl_4]^{2-} + 2e^- \longrightarrow Pt + 4Cl^-$	+0.755
$[PtCl_6]^{2-} + 2e^- \xrightarrow{PtCl_4} {}^{2-} + 2Cl^-$	+0.68
$Pu^3 + 3e^- \longrightarrow Pu$	−2.03
$Ra^{2+} + 2e^- \longrightarrow Ra$	−2.92
$Rb^+ + e^- \longrightarrow Rb$	−2.98
$[RhCl_6]^{3-} + 3e^- \longrightarrow Rh + 6Cl^-$	+0.44
$S + 2e^- \longrightarrow S^{2-}$	−0.47627
$S + 2H^+ + 2e^- \longrightarrow H_2S$	+0.142
$Sc^{3+} + 3e^- \longrightarrow Sc$	−2.09
$Se + 2H^+ + 2e^- \longrightarrow H_2Se$	−0.399
$[SiF_6]^{2-} + 4e^- \longrightarrow Si + 6F^-$	−1.2
$SiO_3^{2-} + 3H_2O + 4e^- \longrightarrow Si + 6OH^-$	−1.697
$SiO_2 + 4H^+ + 4e^- \longrightarrow Si + 2H_2O$	−0.86
$Sm^{3+} + 3e^- \longrightarrow Sm$	−2.304
$Sn^{4+} + 2e^- \longrightarrow Sn^{2+}$	+0.151
$Sn^{2+} + 2e^- \longrightarrow Sn$	−0.1375
$[SnF_6]^{2-} + 4e^- \longrightarrow Sn + 6F^-$	−0.25
$SnS + 2e^- \longrightarrow Sn + S^{2-}$	−0.94
$Sr^{2+} + 2e^- \longrightarrow Sr$	−2.89
$TeO_2 + 4H^+ + 4e^- \longrightarrow Te + 2H_2O$	+0.593
$Th^{4+} + 4e^- \longrightarrow Th$	−1.90
$Ti^{2+} + 2e^- \longrightarrow Ti$	−1.630
$U^{3+} + 3e^- \longrightarrow U$	−1.79
$V^{2+} + 2e^- \longrightarrow V$	−1.19
$Y^{3+} + 3e^- \longrightarrow Y$	−2.37

Standard Electrode (Half-Cell) Potentials	
Half-Reaction	$E°$ (V)
$Zn^{2+} + 2\,e^- \longrightarrow Zn$	−0.7618
$[Zn(CN)_4]^{2-} + 2\,e^- \longrightarrow Zn + 4\,CN^-$	−1.26
$[Zn(NH_3)_4]^{2+} + 2\,e^- \longrightarrow Zn + 4\,NH_3$	−1.04
$Zn(OH)_2 + 2\,e^- \longrightarrow Zn + 2\,OH^-$	−1.245
$[Zn(OH)_4]^2 + 2\,e^- \longrightarrow Zn + 4\,OH^-$	−1.199
$ZnS + 2\,e^- \longrightarrow Zn + S^{2-}$	−1.40
$Zr^4 + 4\,e^- \longrightarrow Zr$	−1.539

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13.2.11: Half-Lives for Several Radioactive Isotopes

Half-Lives for Several Radioactive Isotopes

Isotope	Half-Life[1]	Type of Emission[2]	Isotope	Half-Life[3]	Type of Emission[4]
$^{14}_{6}\mathrm{C}$	5730 y	(β^-)	$^{210}_{83}\mathrm{Bi}$	5.01 d	(β^-)
msubsup	9.97 m	(β^+)	msubsup	60.55 m	(α or msupmsup)
$^{15}_{9}\mathrm{F}$	4.1×10^{-22} s	(p)	msubsup	138.4 d	(α)
msubsup	15.00 h	(β^-)	msubsup	3×10^{-7} s	(α)
$^{32}_{15}\mathrm{P}$	14.29 d	(β^-)	$^{216}_{84}\mathrm{Po}$	0.15 s	(α)
msubsup	1.27×10^{9} y	(β or $E.C.$)	msubsup	3.05 m	(α)
msubsup	0.08 s	(β^+)	msubsup	1.0×10^{-4} s	(α)
msubsup	2.6×10^{6} y	(β^-)	msubsup	1.6 s	(α)
msubsup	5.27 y	(β^-)	msubsup	55.6 s	(α)
$^{87}_{37}\mathrm{Rb}$	4.7×10^{10} y	(β^-)	msubsup	3.82 d	(α)
$^{90}_{38}\mathrm{Sr}$	29 y	(β^-)	$^{224}_{88}\mathrm{Ra}$	3.66 d	(α)
$^{115}_{49}\mathrm{In}$	5.1×10^{15} y	(β^-)	$^{226}_{88}\mathrm{Ra}$	1600 y	(α)
$^{131}_{53}\mathrm{I}$	8.040 d	(β^-)	$^{228}_{88}\mathrm{Ra}$	5.75 y	(β^-)
$^{142}_{58}\mathrm{Ce}$	5×10^{15} y	(α)	$^{228}_{89}\mathrm{Ac}$	6.13 h	(β^-)
$^{208}_{81}\mathrm{Tl}$	3.07 m	(β^-)	$^{228}_{90}\mathrm{Th}$	1.913 y	(α)
$^{210}_{82}\mathrm{Pb}$	22.3 y	(β^-)	$^{232}_{90}\mathrm{Th}$	1.4×10^{10} y	(α)
$^{212}_{82}\mathrm{Pb}$	10.6 h	(β^-)	$^{233}_{90}\mathrm{Th}$	22 m	(β^-)
$^{214}_{82}\mathrm{Pb}$	26.8 m	(β^-)	$^{234}_{90}\mathrm{Th}$	**24.10 d**	(β^-)
$^{206}_{83}\mathrm{Bi}$	6.243 d	($E.C.$)	$^{233}_{91}\mathrm{Pa}$	27 d	(β^-)
$^{233}_{92}\mathrm{U}$	1.59×10^{5} y	(α)	$^{242}_{96}\mathrm{Cm}$	162.8 d	(α)
$^{234}_{92}\mathrm{U}$	2.45×10^{5} y	(α)	$^{243}_{97}\mathrm{Bk}$	4.5 h	(α or $E.C.$)
$^{235}_{92}\mathrm{U}$	7.03×10^{8} y	(α)	$^{253}_{99}\mathrm{Es}$	20.47 d	(α)
$^{238}_{92}\mathrm{U}$	4.47×10^{9} y	(α)	$^{254}_{100}\mathrm{Fm}$	3.24 h	(α or $S.F.$)
$^{239}_{92}\mathrm{U}$	23.54 m	(β^-)	$^{255}_{100}\mathrm{Fm}$	20.1 h	(α)
$^{239}_{93}\mathrm{Np}$	2.3 d	(β^-)	$^{256}_{101}\mathrm{Md}$	76 m	(α or $E.C.$)
$^{239}_{94}\mathrm{Pu}$	2.407×10^{4} y	(α)	$^{254}_{102}\mathrm{No}$	55 s	(α)
$^{239}_{94}\mathrm{Pu}$	6.54×10^{3} y	(α)	$^{257}_{103}\mathrm{Lr}$	0.65 s	(α)
$^{241}_{94}\mathrm{Pu}$	14.4 y	(α or msupmsup)	$^{260}_{105}\mathrm{Ha}$	1.5 s	(α or $S.F.$)
$^{241}_{95}\mathrm{Am}$	432.2 y	(α)	$^{263}_{106}\mathrm{Sg}$	0.8 s	(α or $S.F.$)

Table M1

Footnotes

- 1y = years, d = days, h = hours, m = minutes, s = seconds
- 2*E.C.* = electron capture, *S.F.* = Spontaneous fission
- 3y = years, d = days, h = hours, m = minutes, s = seconds
- 4*E.C.* = electron capture, *S.F.* = Spontaneous fission

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